A COURSE

IN

THE ELEMENTARY PRINCIPLES

OF

CHEMISTRY

FOR SECONDARY SCHOOLS

FOURTH EDITION

BY

B. W. McFarland, C.E., Ph.D.

Instructor in Chemistry

IN THE

NEW HAVEN HIGH SCHOOL

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A

THIRTY-FIVE WEEKS' COURSE

IN

THE ELEMENTARY PRINCIPLES OF CHEMISTRY

FOR SECONDARY SCHOOLS.

TO BE USED IN CONNECTION WITH ANY GOOD DESCRIPTIVE CHEMISTRY.

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PART I. LABORATORY WORK.

PREFACE.

In publishing this book the author is primarily influenced by a desire to place in the hands of his own students a set of directions for their laboratory work, combined with a concise statement of those definitions, principles, and explanations which he considers essential to an intelligent understanding of the subject.

The course represents what is being done in this laboratory at present and what has been done for several years. It is in reality the outcome of a good many years of experience in teaching the subject from a variety of text books without any printed laboratory guide whatever, combined with the necessity of obtaining practical results in a very limited time. The available time for the subject in this school consists of thirty-five laboratory periods of eighty minutes each and seventy recitation periods of forty minutes each.

In developing the subject it has been the constant aim to teach general principles and to eliminate as far as possible the burden resulting from the mere memorizing of a large quantity of ill-assorted facts. With this in view the laboratory work is so arranged that for the first three months the student meets only such chemical changes as are typical of the fundamental reactions of chemistry. With the exception of oxygen, hydrogen, and water, no substances are studied during this time that require much descriptive work in recitation. The extra time is spent in constant drill on fundamental definitions, the meaning and use of symbols, formulas and equations, and on the subject of chemical equivalence. As a result of this drill the student becomes perfectly familiar with the language of chemistry and has no further serious trouble in remembering, classifying and expressing chemical facts.

Beginning with the subject of chlorine, and continuing through bromine and iodine, the general subject of oxidation is taken up, and the use of many oxidizing agents is studied.

At each following laboratory exercise one or more substances are studied. The results of the laboratory work are discussed and explained in the next recitation.

The student is required to take careful original notes in a suitable note book of all results obtained in the laboratory and to write a composition in the book on each substance studied, according to a specified outline. The material for this he gets partly from his laboratory work and partly from his text book or other book of reference.

At the second recitation following the laboratory work the substances studied are described according to the outline.

Exercises in performing examples and in writing equations are constantly given to a few students during each recitation; and whenever opportunity offers, whole lessons are assigned from the review work indicated in Tables I, II and III in Part IV.

The ideal condition in teaching chemistry would be to have one teacher to each student. A student under those conditions would be sure to like the subject and do well in it, for the course would be exactly adapted to his needs. To adapt the teaching to a large class is, however, the actual condition which confronts the instructor. He must neither overwork the slow student, nor allow the bright one to be idle and so to feel that he is getting little from his subject. He must also see that the average student, who is very willing to work but unwilling to search long for the ideas needed in his work, shall find those ideas at hand. These problems, it is hoped, will be solved, to a certain extent, by giving to the student, both in his laboratory work and in his outside study, a book which will partially take the place of an instructor by his side.

On account of the short time available, the quantity of material used in the experiments is extremely small. In order to specify these quantities, when referring to dry reagents, a particular small double horn spoon, more fully described in Part IV, has been adopted in this laboratory and is referred to as small spoonfuls "s.s.," or large spoonfuls, "l.s."

All reference numbers refer to paragraphs.

The use of a pipe system for the distribution of gases to the students' tables need in no way interfere with the use of the book in laboratories not so equipped, for the experiments in which it is used are rather few and could be readily omitted without damaging the course; or other experiments could be substituted for them.

No apparatus is used that need take more than one minute to put in place. All experiments that might result in accidents destructive to apparatus or material have been avoided, as well as those that frequently give unsatisfactory results in the hands of students.

The course assumes that the student has had a good high school course in physics, and is intended to be used in the fourth or senior year. It is, however, used to some extent by second and third year classes in this school, and the results are satisfactory.

In conclusion the author wishes to express his thanks to Dr. P. T. Walden and Dr. H. W. Foote, both of the Sheffield Scientific School, for many valuable suggestions and kindly criticisms.

It would give the author pleasure to correspond with teachers who may use the book, and to receive from them, or others interested, suggestions or criticisms.

TO THE TEACHER.

The following course may be considered to be divided into three distinct parts. The first part has to do entirely with teaching the language of chemistry, i. e. the meanings and uses of symbols, formulas and equations as well as the fundamental reactions, and lasts about twelve or fourteen weeks. In addition to the first twelve laboratory exercises the student is thoroughly drilled on everything in Part II to art. 65. Especial stress should be laid upon the following topics:

1. Symbols and names of the common elements in groups.
2. Full meaning of a formula of a compound.
3. Combining powers or valence of the different elements.
4. Formulas of binary compounds.
5. How many grams of one element would be chemically equivalent to a certain number of grams of another element?
6. Full meaning of a chemical equation.
7. Equations showing the formation of binary compounds from the elements.
8. Symbols and names of the radicals.
9. Formulas and names of the compounds of all the positive radicals with all the negative radicals.
10. Seven general methods for the preparation of salts.

This can be accomplished to the entire satisfaction of both pupil and teacher in about twenty-two recitations.

As a result of the above drill the student becomes so familiar with the meaning and use of formulas and equations that during the rest of the course he experiences no trouble whatever in recording, remembering and explaining the chemical changes that come under his observation. It also leads to an almost unconscious knowledge of the general subject of chemical equivalence which is of the utmost value.

The second part of the course lasts about twelve weeks. During this time the non-metals are especially studied both in the laboratory and recitation room. At the first recitation following a laboratory exercise the experiments are thoroughly discussed and explained, the student being expected to state what he did and what he observed and to explain what happened, as far as possible, representing all changes by equations.

This recitation is as much an instruction exercise as it is a recitation, but the student receives full credit for having made the observations that might reasonably be expected from the experimental work done.

During this portion of the course the subject of oxidation in the wet way is thoroughly taken up and instruction is given in forming equations representing oxidation reactions.

At the second recitation following a laboratory exercise, the subject of the laboratory work, when it is a substance, is recited according to a specified outline which the student fills out partly from the laboratory work and partly from the text book. This results in much repetition of certain ideas of great importance which are otherwise apt to be neglected. In this way the text book becomes a reference book and is used intelligently, the burden of remembering mere print being entirely eliminated.

Following this, the student is required to write an article in the form of a report on the substance studied, covering the main properties and describing and explaining what was done with it in the laboratory.

This provides an immediate review of what was learned in the two recitations and requires an additional reading of the matter in the text book; the whole constituting a very desirable preparation for the next laboratory exercise at which the composition is due.

In the last part of the course the laboratory study of the metals is completed, and during the recitations the special review exercises indicated in Tables I, II and III are taken up. The exercise indicated in Tables II and III is of the greatest value in giving a review of the whole subject. In this exercise the student is expected to represent by equations the transformation of natural elements and compounds into common compounds and elements.

In this work he is not allowed to make use of the formula of any substance whose formation he has not already shown from the natural substances. This, again, is an exercise that gives great repetition of the very points most needed and brings out strongly the relative importance of many processes and principles.

During the last two or three weeks all recitations are stopped and the student is required to spend all the time in the laboratory working on the identification of common chemical compounds. In this work he shows a remarkable amount of interest for the end of the year.

As a review of the laboratory work of the year the exercise is unexcelled.

Examples should be constantly given. At first only those involving weights, then those involving volumes of gases, finally those involving volumes, specific gravities and percentage content of liquids.

PART I.

LABORATORY WORK.

INTRODUCTION.

Personal Equipment.

Each student should provide himself with an apron, a towel, a piece of cloth, a sponge, and a piece of soap.

The apron may be best made of denim, perfectly plain with a fixed strap around the neck and a button strap around the waist. It should be long enough to reach within four inches of the floor.

Individual Laboratory Equipment.

Learn the names of these things the first day in the laboratory and the places where they belong.

To be kept in drawer.

Long compartment.	¼ lb. glass tubing. 1 funnel tube. 1 combustion spoon. 1 glass bend, 45 cm. long. 3 pieces of copper wire.
Short compartment.	1 piece of rubber tubing. 1 file. 1 pair of forceps. 1 glass stirring rod. 1 iron wire loop.
First square.	8 squares of glass. 3 rubber connectors.
Second square.	1 porcelain mortar with pestle. 3 No. 8, 2-hole rubber stoppers. 1 No. 3, 1-hole rubber stopper.

Large section.

- 3 short right angle bends, 8 cm. x 8 cm.
- 3 medium right angle bends, 8 cm. x 15 cm.
- 1 3″ Royal Berlin porcelain evaporating dish.

To be kept in locker.

On the shelf.

- 1 No. 2 beaker.
- 1 8 oz. flask.
- 1 funnel.

On the floor.

- 6 5 oz. bottles.
- 2 500 c.c. bottles.
- 1 test tube rack.
- 12 test tubes in rack, upside down.
- 1 test tube holder on rack.

Equipment for General Use.

To be kept in locker.

- 1 water pan.
- 1 glass shelf for water pan.
- 1 iron lamp stand, two rings and one clamp.
- 1 filter flask with rubber tubing.
- 1 porcelain sieve funnel with rubber stopper.
- 1 1000 c.c. bottle.

To be kept in drawer.

- 1 centimeter scale.
- 1 bulb test tube.
- 1 package of filter paper.
- 1 bundle of splints.
- Disks of filter paper for porcelain funnel.
- 1 wire gauze with asbestos center.
- 1 iron crucible
- 1 iron plate.
- 1 piece of cobalt glass.
- 1 thermometer.
- 1 package of common wrapping paper.

To be kept on table.

- 1 50 c.c. graduate, or measuring glass.
- 8 reagent bottles.
- 1 incandescent lamp.
- 1 litmus paper bottle.
- 2 Bunsen burners with tubing.
- 1 test tube cleaner.

STUDENTS ARE EXPECTED TO KEEP THEIR SUPPLY OF APPARATUS COMPLETE. THIS CAN BE DONE BY PAYING AT ONCE FOR THOSE THINGS THAT ARE BROKEN OR LOST THROUGH CARELESSNESS, AND RECEIVING NEW MATERIAL.

Care of the Table and Lockers.

SPONGE OFF THE TABLE AFTER EVERY LABORATORY EXERCISE WITH A MOIST SPONGE, WHETHER ANYTHING HAS BEEN SPILLED ON IT OR NOT, AND WIPE IT DRY WITH A CLOTH.

Keep the reagent bottles in the proper order, viz., from left to right—sulphuric acid, conc., dil.; nitric acid, conc., dil.; hydrochloric acid, conc., dil.; ammonium hydroxide and sodium hydroxide.

Wash all the reagent bottles and sponge off the shelves once a week.

Keep all of your own apparatus clean and in the right place.

DON'T put the general apparatus in your private locker.

SOME LABORATORY DON'TS.

Don't throw solids in the sink.

Don't boil sulphuric acid.

Don't pour water into hot sulphuric acid.

Don't aim a test tube at your neighbor. Point it to the right, down the middle of the table.

Don't put **anything** into the reagent bottles, not even the reagent itself.

Don't mix chemicals on your own account. Students who are found doing this will be dismissed from the laboratory.

Don't use large quantities where small quantities will do just as well.

Don't depend entirely on your own observation for your facts; verify them by reading up the subject in the various reference books.

Don't pay any attention to the way your neighbor is doing his work.

REMEMBER.

Neatness is of the first importance.

You will be marked on the way you keep your desk and locker.

To follow directions implicitly, but try to understand the reasons for them.

Where the reference books are kept.

You are responsible for the results whether you do the work or not.

To bring your note book and text book to all exercises, both recitations and laboratory work.

You are at liberty to use the laboratory at all times, when it is not occupied by a regular class.

To keep your eyes open.

To read the signs.

To learn the names of the common minerals and chemicals in the wall cases.

THE NOTE BOOK.

All writing in the note book is to be in lead pencil. A hard pencil should be kept for this purpose.

Every laboratory exercise should appear in the book in three parts, as follows:

Part I. Is to contain whatever the instructor may have to dictate before the class goes to the tables. It will contain definitions and principles that are needed in the day's work.

Part II. Is the most important part of the book. It is to contain the **original** notes of the student on the experiment in question.

The notes must be short, on separate lines, contain a brief reference to what was done to produce the result, and must be numbered or lettered to correspond with the directions.

Part III. Is to be a composition on the day's laboratory work. The particular subject will be assigned by the instructor each day.

If the subject of the composition is a substance, the outline for it will be as follows:

GENERAL OUTLINE TO BE FOLLOWED IN RECITATION AND IN WRITING COMPOSITIONS.

Appearance and state.	
Occurrence in the free condition.	Where ? In what quantity ?
Occurrence in the combined condition.	Where ? As what ?
Physical Properties.	Can it be liquefied, solidified, or gasified, and under what conditions ? Will it dissolve in water or in other liquids ? To what extent ? At what temperature does it melt or boil ? How does its volume change with a change of temperature ? **101** How does its volume vary with the pressure ? What is its specific gravity ? If it is a gas, what does a liter weigh ?
Chemical Properties.	Will it burn or support combustion ? Can it be decomposed ? Does it combine with oxygen ? What oxides does it form ? Are these oxides acidic or basic ? Does it form any compounds with hydrogen? **What experiments might be done with it ?** **Describe them in detail.** **Give all equations involved.**
Preparation.	The common way it is made in the laboratory. The common way it is manufactured. Two other ways. Give all equations.
Common compounds not already described.	The chemical name and formula. The common name and the mineral name.
Uses.	In manufacturing. In medicine.

The common test for the substance, if it is an element.

NOTE. It is not expected that answers will be found for every one of these questions for any one substance. Whatever facts are learned either from the laboratory work or from the text book are to be fitted into the outline.

SPECIMEN EXERCISE.

July 24. Laboratory Exercise No. 49.

EXPERIMENTS WITH MANGANESE.

Part I.

1. Heat a small quantity of manganese dioxide in a sealed tube. What gas is given off, and what is left in the tube? **71.**

2. Heat ¼ g. of manganese dioxide in a t.t. with 5 c.c. conc. HCl. What gas is given off? What kind of a reaction was this? **75.**

3. Shake up some manganese dioxide with some strong sulphurous acid water.

4. Add a small piece of sodium nitrite to a tube containing ¼ g. of manganese dioxide and 5 c.c. of dilute nitric acid. Heat to boiling, then add ½ g. of red lead. **78.** What compound of manganese is formed? Write all reactions.

5. Fuse 1 g. of manganese dioxide with 1 g. of potassium carbonate and a little potassium nitrate in iron crucible. Dissolve out the product with boiling water, filter and pass carbon dioxide into the solution.

Part II.

1. a. Oxygen given off, shown by spark.
   b. Red powder left in tube, Mn_3O_4.

2. Chlorine given off and manganous chloride formed in solution. This is an oxidation reaction.

3. The sulphurous acid is oxidized to sulphuric acid, and the manganese dioxide is decolorized.

4. The sodium nitrite with the nitric acid gives off fumes of nitrogen trioxide, red; these form nitrous acid with the water, and this is oxidized by the manganese dioxide to nitric acid, the manganese being changed into the nitrate. The red

lead oxidizes the manganese nitrate to permanganic acid, $HMnO_4$, which is purple in solution.

5. When the two substances are fused the compound is green. This contains potassium manganate.

When CO_2 is passed into a solution of this, it is changed to a purple solution.

Part III.

Manganese is a metal resembling iron. It does not occur in the free condition, but is very abundant in the form of oxides, the most common one being manganese dioxide or pyrolusite. It is a hard, brittle metal having a grey color and a high lustre which it soon loses in moist air. In almost all of its physical properties it resembles iron, with which it is very commonly associated, both in the ore and in the finished metal. It forms several oxides, of which the lowest are basic, forming salts with acids, and the highest acidic, forming manganic and permanganic acids.

Appearance
Occurrence
combine

Physical
properti

Chemical
properti

When manganese dioxide is heated, it gives up part of its oxygen and becomes a brown substance having the formula Mn_3O_4. If manganese dioxide be heated with hydrochloric acid, some of its oxygen unites with the hydrogen of the acid, thereby setting free chlorine. This is the common method of preparing chlorine in the laboratory. The dioxide itself is reduced to the basic manganous oxide, MnO, and this unites with some of the hydrochloric acid present and forms manganous chloride. Sulphurous acid is oxidized to sulphuric acid by manganese dioxide, the manganese going into solution, partly as manganese dithionate and partly as manganese sulphate. If sodium nitrite be treated with nitric acid, it will be changed to sodium nitrate, and nitrous acid will be set free. If manganese dioxide be present, the nitrous acid will reduce it to manganous oxide, becoming changed itself to nitric acid. The manganous oxide unites with the nitric acid and becomes manganous nitrate. If some red lead be added to this solution, boiling hot, the manganese will be oxidized to permanganic acid, which colors the liquid purple. When manganese dioxide is fused with potassium carbonate, it is changed to potassium manganate. This is green in color. And if this

be dissolved in water, filtered, and carbon dioxide passed in, it is changed to potassium permanganate. Equations for these reactions are as follows:

$$3MnO_2 \text{ heated} = Mn_3O_4 + O_2.$$
$$MnO_2 + 4HCl = MnCl_2 + Cl_2 + 2H_2O.$$
$$MnO_2 + H_2SO_3 = MnSO_4 + H_2O.$$
$$MnO_2 + 2H_2SO_3 = MnS_2O_6 + 2H_2O.$$
$$NaNO_2 + HNO_3 = NaNO_3 + HNO_2.$$
$$HNO_2 + MnO_2 = MnO + HNO_3.$$
$$MnO + 2HNO_3 = Mn(NO_3)_2 + H_2O.$$
$$26HNO_3 + 2Mn(NO_3)_2 + 5Pb_3O_4 = 2HMnO_4 + 15Pb(NO_3)_2 + 12H_2O.$$

paration. Manganese can be prepared by reducing the oxide. This can be done with carbon at a high temperature. Or the dioxide may be mixed with the proper quantity of aluminum and the mixture ignited. The aluminum unites with the oxygen and the manganese is set free in the liquid condition. A very high temperature is produced by the reaction.

$$MnO_2 + C = CO_2 + Mn.$$
$$3MnO_2 + 4Al = 2Al_2O_3 + 3Mn.$$

es. The metal is used to alloy with iron and steel.

Manganese dioxide or black oxide of manganese, MnO_2, is used in the laboratory to prepare chlorine.

nmon compounds. Potassium permanganate is a common laboratory oxidizing agent.

t. The element is most easily recognized by melting a little of the material to be tested, with sodium carbonate and potassium nitrate, on platinum foil. If any manganese is present, the green color of sodium manganate will be seen.

LABORATORY WORK.

Abbreviations used in the directions.

c.c., cubic centimeters.

g., grams.

s.s., a small horn spoonful. The spoon holds about 1½ g. of mercuric oxide.

l.s., a larger horn spoonful. The spoon holds about 9 g. of mercuric oxide.

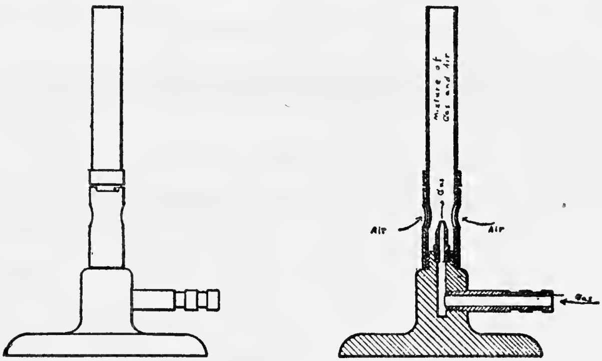
t.t., test tube.

ppt., precipitate.

LABORATORY EXERCISE NO. 1.

Construction and Operation of the Bunsen Burner.

1. Note the general construction of it. 2. Take it apart and make a drawing of each piece. 3. Describe how to light it. 4. Four ways in which, sooner or later, it will be found burn-



Side view. Fig. 1. Cross section.

ing. Only one correct way. 5. How to correct it when it is found burning at the base. 6. Make a drawing of the Bunsen flame. Note two parts to it. 7. Heat a piece of copper wire in and about the flame. **68.** From this decide what is the hottest point in the flame. Is the flame hollow in the middle?

LABORATORY EXERCISE NO. 2.

Working with Glass.

1. To cut glass tubing, make a short sharp scratch across the tube with a file, place the file on the table directly under the scratch, with the scratch uppermost, then bear down gently on the tube so as to break it.

2. To round the edges of a glass tube, place the end of the tube in the hottest part of the flame and rotate it slowly. Hold the other end lower down.

3. To seal the end of a tube, hold the end in the hottest part of the flame, slowly rotating it until it seals up.

4. To make a bulb tube, seal the end of a tube as above, then remove it from the flame and blow carefully at once until the bulb is of the required size.

5. Ordinarily, to make sealed tubes, take a piece of tubing that is just twice as long as the desired tube and seal it off in the middle, thereby making two sealed tubes at once.

6. To bend glass tubing, heat it lengthwise in an ordinary gas flame until it is quite soft; then remove it from the flame and bend it quickly to the required curve. **Do not attempt to bend it after the glass begins to harden.**

What new observations have you made during this exercise?

LABORATORY EXERCISE NO. 3.

The Collection of Gases over Water and some Properties of Oxygen.

Materials. Magnesium; oxygen; splints; magnesium oxide.

Apparatus. Water pan; bottles; delivery tube.

Definitions—Acid 55; Alkali 57; Basic oxide 60; Alkaline oxide 62; Acidic oxide 61.

A **delivery tube** is any combination of rubber or glass tubing by means of which a gas can be transferred from one place to another.

A **water pan** is used for the collection of gases over water.

1. Fill several bottles with oxygen. Holding a splint by means of the forceps, ignite it and thrust it down into one of the bottles. **68.**

2. Repeat with magnesium ribbon. **68.** Burn several pieces. Scrape some of the product out on a glass plate and find out whether it reacts acid or alkaline toward moist litmus paper. **62.**

3. Test the action of magnesium oxide in water and in dilute acids. **60.** In which does it dissolve readily? Is it a basic or an acidic oxide?

4. Make very dilute solutions of each of the reagents on the shelf, in test tubes. Test these solutions with litmus paper. Which are acid, and which are alkaline?

Never lay the stopper of a bottle down on the table. Turn the hand over, take the stopper between the first and second fingers, remove it from the bottle, pick up and use the bottle with the same hand, then replace the stopper.

LABORATORY EXERCISE NO. 4.

Some Properties of Metallic Oxides.

Materials. Sodium wire, made before the class; mercuric oxide, HgO; splints; manganese dioxide, MnO_2; red oxide of lead, Pb_3O_4; calcium oxide, CaO; magnesium oxide, MgO.

Apparatus. Iron crucible; lamp stand; funnel; filter paper.

Definitions—Properties of matter. **5.**

Filtrate. The liquid which runs through a filter.

Residue. The material left on the filter.

Note. A **solution is always transparent.** If a liquid is not transparent it probably contains a finely divided solid that remains in suspension and could be filtered off.

How to use the lamp stand. How to make a filter. How to filter. Use of stirring rod.	To be demonstrated by the instructor.

1. Place a small piece of sodium on a piece of wet filter paper. Turn up the edges of the paper so the melted sodium cannot roll off. The white smoke is sodium oxide.

2. Heat a piece of sodium wire in iron crucible on lamp stand until no further change takes place. Let it cool for half a minute, add water, heat to boiling and filter into t.t. Feel the filtrate with the fingers and note its action on litmus paper. What is an alkaline oxide? **62.**

3. Place a small quantity of mercuric oxide in a sealed tube, not more than ½ of a s.s., heat and test the escaping gas with a spark on the end of a splint. What is left in the tube? **71.**

4. Heat manganese dioxide in sealed tube, and test escaping gas for oxygen. **71.**

5. Repeat, using red lead. **78-7.**

6. Test the solubility of each of the oxides mentioned at the beginning of this exercise, in the dilute acids, and tabulate the results. Which are alkaline oxides? **62.** Which are basic? Which one is not a basic oxide? Do all oxides give oxygen when heated?

LABORATORY EXERCISE NO. 5.

Preparation and Properties of Oxygen and Oxides of Non-metals.

Materials. Manganese dioxide, MnO_2; oxygen mixture (½ MnO_2 and ½ $KClO_3$); charcoal, C; sulphur, S; phosphorus, P **176**; arsenic trioxide, As_2O_3.

Apparatus. Water pan; lamp stand; bottles; combustion spoon; iron plate; delivery tube.

Definitions—Acidic oxide 61; Acid anhydride 63.

1. Fit a t.t. with a rubber stopper and a delivery tube. Place in the tube about one inch of the oxygen mixture. Fix the

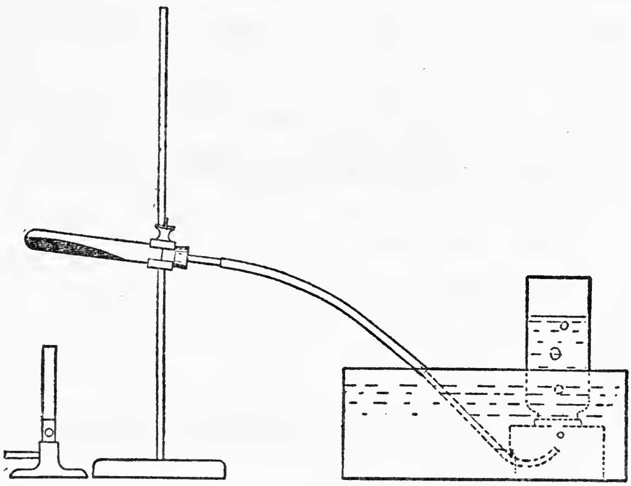


Fig. 2.

tube in the clamp of the lamp stand, in a nearly horizontal position as in fig. 2. The mouth of the tube should be a little lower than the heated portion. Heat the mixture, being

careful not to melt the tube, and collect the gas in bottles in the water pan. **71.** If this does not give enough gas to fill all the bottles, fill the rest from the supply pipe.

2. Balance a small piece of charcoal on the combustion spoon, ignite it and lower it into a bottle of oxygen. **68.** When it will burn no longer add a small quantity of water to the bottle, cover it with the palm of the hand tightly, and shake the water about in the bottle. What kind of oxides do most non-metals form? **61-63.** Test this water with litmus.

3. Pick up a small bead of phosphorus with the forceps and dry it by touching it to filter paper. Place it in the cold dry combustion spoon. Ignite it in the hood and lower it in a bottle of oxygen carried there for the purpose. What is the white smoke? What becomes of it? Add water to the bottle, shake and test with litmus paper.

4. Ignite some sulphur in combustion spoon and lower it into a bottle of oxygen. Add water and shake. Test the water with litmus paper.

5. Place a dry piece of phosphorus on a dry iron plate; ignite it and cover it at once with a large dry bottle. Let it stand until the fumes have settled. These fumes are phosphorus pentoxide. Scrape some of the oxide out on a piece of glass. Breathe on it. Add one drop of water to it. Test the product with litmus paper.

6. Test arsenic trioxide with litmus paper. See if it will dissolve in an alkali. What kind of an oxide is it?

LABORATORY EXERCISE NO. 6.

Properties and Preparation of Hydrogen.

Materials. Granulated zinc; splints.

Apparatus. Test tube; one-hole rubber stopper; glass bends; rubber connectors; water pan; bottles; glass plates.

1. Making use of the water pan, fill all the bottles with hydrogen from the supply pipe. Keep them inverted and covered with glass plates.

2. Does hydrogen burn? Attach a delivery tube to the supply pipe, fit the end with a glass tip, turn on a slow stream of gas and attempt to light it. What color has the hydrogen flame?

3. Does it support combustion? Place one of the covered bottles upside down on a ring of the lamp stand, which will support the bottle when the glass plate is removed. Fasten a splint on a wire. Now remove the glass plate and support the bottle on the ring, covering it again at once with the glass plate held in the left hand. Ignite the splint and quickly plunge it up into the bottle the instant the plate is removed. Withdraw it slowly. What do you see?

4. Does this gas dissolve in water? Add about an inch of water to a bottle of hydrogen, cover the bottle tightly with the palm of the hand, and shake it thoroughly. If the gas is soluble in water, there will be a slight suction on the hand.

5. Is hydrogen lighter than air? Fill two bottles with the gas. Remove the cover of one and leave it standing right side up for one minute. Try to ignite the gas in the bottle at the end of that time. Uncover another bottle, but keep it upside down for one minute (this can best be done by supporting it on the lamp stand). Then ignite the gas. Obtain a clay pipe, and blow soap bubbles with hydrogen from the supply pipe. Have a bottle full of hydrogen upside down on a bottle of the same size full of air. Invert the two bottles and after three minutes test them both to see which contains the hydrogen.

6. Is a mixture of hydrogen and air explosive? Mix hydrogen and air in various proportions and attempt to ignite the mixture. If it is explosive, ascertain what the most explosive mixture is.

7. What is formed when hydrogen burns? Insert a small jet of burning hydrogen in a dry bottle. Is any moisture noticeable?

8. Hydrogen does not occur free. It must be separated from some of its compounds. How can it be obtained from water? It can most easily be obtained from acids in the laboratory by putting some metal with them. **64a.** Pour some dilute sulphuric acid on some granulated zinc in a test tube. How could you collect this gas? Fit the tube with a rubber stopper and a delivery tube and collect several bottles of the gas. Prove that this gas is the same as that with which you have just been experimenting and which is known to be hydrogen. In how many ways could you prove it to be the same?

In case it is desired to generate a larger quantity of hydrogen than can well be obtained from a test tube, the apparatus indicated in fig. 3 should be used. The funnel tube is passed through a ring of the lamp stand simply to prevent the apparatus from being knocked over. It goes to the bottom of the 250 c.c. bottle, A, containing one or two inches of granulated zinc. A short bend comes out of the two-holed rubber stopper in A and is connected by a rubber connector to a medium bend that passes through another rubber stopper and goes to the bottom of the wash bottle, B, which is *three quarters full*

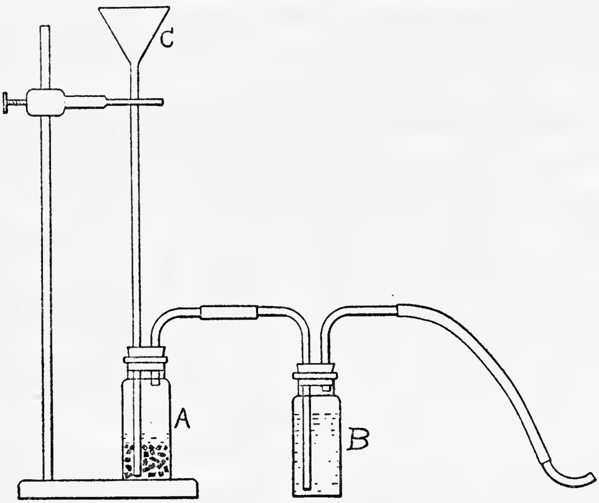


Fig. 3.

of water. Enough water is now poured through the funnel tube to cover the zinc in A. Conc. sulphuric acid is then added little by little, **very cautiously,** until the desired reaction is obtained. Sometimes the zinc is too pure to react properly with pure acid, in which case the action will start readily if a little copper sulphate solution be added.

Never ignite a jet of hydrogen coming from a generator in operation.

9. When the zinc comes in contact with the sulphuric acid, each atom of zinc displaces two atoms of hydrogen from the acid and forms a compound called zinc sulphate. **64a.** Cover

the bottom of the beaker with granulated zinc. Cover the zinc with dilute sulphuric acid. Put this away in your locker until the next time. What do you find besides some undissolved zinc? **64a.**

10. Fasten a glass jet-tube 20 cm. long, drawn out to a small hole, vertically in the clamp of the lamp stand, connect this with the supply pipe, turn on a small stream of hydrogen, and ignite the jet. Put slowly down over this jet, glass tubes of various sizes. The noises are called hydrogen tones.

LABORATORY EXERCISE NO. 7.

Electrolysis.

Materials. Sodium chloride; potassium sulphate; litmus solution.

Apparatus. Voltameter.

Definitions—

Electrolysis, the process of decomposing a dissolved or melted substance by means of the electric current.

Electrolyte. A substance which, when dissolved or melted, will allow the passage of the electric current.

Electrodes. The terminals immersed in a liquid by means of which the current enters and leaves.

Positive electrode. The one on which the oxygen appears.

Negative electrode. The one toward which the current travels and on which the hydrogen appears.

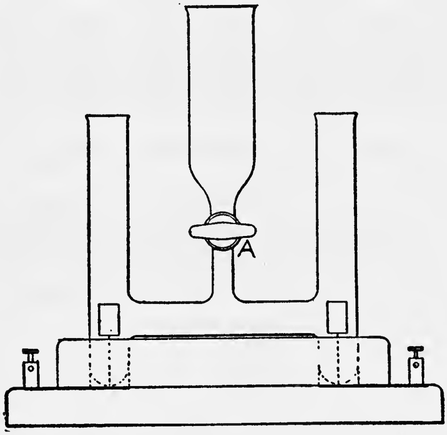
Questions to be answered by experiment.

1. Will distilled water allow the current to pass through it?
2. Will hydrant water allow the current to pass through it?
3. What effect has a small quantity of sulphuric acid on the conductivity of distilled water?
4. Will the current pass through a very dilute solution of sodium hydroxide? **124.**
5. How does the presence of salt of any kind affect the passage of the current? **124.**
6. What will be the effect of passing a current through a very dilute solution of potassium sulphate colored with litmus solution? **124.**
7. What classes of substances, when dissolved, allow the passage of the current through water?

1. Clean the voltameter by taking the glass part out of the base and running hydrant water through it, then rinse with distilled water. Fill it with distilled water and replace it in the support. Turn on the current. (110 volt current through one 32 c.p. lamp.) Judge of the amount of current passing by the rate at which hydrogen is generated.

2. Repeat exactly, using hydrant water. Note that in each case very little current passes but that much more gets through the hydrant water than through the distilled water.

3. Make a mixture of 100 c.c. of water and 5 c.c. of conc. sulphuric acid. Pour this into the apparatus until the level of



Voltameter.

Fig. 4.

the liquid comes nearly to the top of the side tubes. Have the stop-cock at A open and fit solid rubber stoppers into the top of the side tubes so there will be no air inclosed. Then turn on the current. Note the time. Let the current run until one tube is just filled with a gas. Note the time. Shut off the current and close the stop-cock at A. Remove one of the stoppers and test the gas at once with a flame on a splint. Replace the stopper and test the gas on the other side. Which was hydrogen? How much hydrogen was there compared to oxygen? Find what volume of hydrogen was formed in the measured time. The current you had was about .8 of an

ampere. Find how many c.c. of hydrogen one ampere would give in one second. One ampere should give exactly .0000103 g. of hydrogen in one second.

4. Test solutions of sodium hydroxide and sodium chloride in the same way. What effect have acids, bases and salts in solution on the conductivity of water? **118.** Why did hydrant water probably conduct the current better than distilled water? **124.**

5. Make a solution of pure potassium sulphate and color it by the addition of some litmus solution. **124.** The voltameter must be thoroughly washed before adding the liquid. Turn on the current and note the effect. When oxygen salts of sodium or potassium are electrolyzed an acid is formed on one side and an alkali on the other.

LABORATORY EXERCISE NO. 8.

The Reduction of Metallic Oxides and the Preparation and Purification of Water.

Materials. Copper scale, a mixture of cupric oxide, CuO, and cuprous oxide, Cu_2O; lead oxide or litharge, PbO; potassium permanganate solution, $KMnO_4$; crystallized copper sulphate or blue vitriol, $CuSO_4 + 5H_2O$; copper gauze.

Apparatus. Lamp stand; wire gauze; flask; No. 8 two-hole rubber stopper; delivery tube; funnel; filter paper; sealed tubes.

Definitions—Reduction. 67. Reducing agent. 80.

A **precipitate** is a solid that separates from solution. In the laboratory it is commonly formed by mixing two transparent solutions. When this is done, if the mixture is not perfectly transparent, a precipitate has been formed.

Crystalline. A substance is said to be crystalline when it possesses naturally formed faces that reflect light.

Amorphous. Not crystalline.

Water of crystallization. Many substances when crystallizing from a water solution, combine chemically with a definite number of molecules of water. This water is necessary for the crystalline form and color of the substance.

Efflorescence. Many substances containing water of crystallization give up part or all of this water to the air when exposed, thereby losing their crystalline form and color, and become amorphous powders. Such substances are said to effloresce.

Distillation. The process of boiling a substance and condensing the vapor.

Distillate. The liquid that comes out of a condenser, condensed vapor.

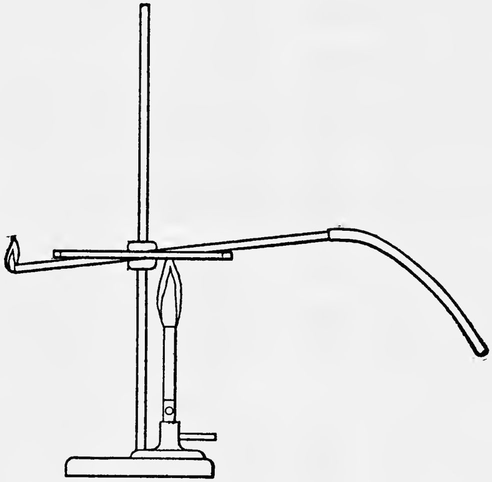


Fig. 5.

1. Hydrogen has a very great attraction for oxygen. It will not only combine with it directly, but it will extract it from other compounds and combine with it to form water. **80-81.** Push a small piece of copper gauze, by means of the handle of the combustion spoon, a little more than half way through a glass tube about eighteen centimeters long. Add to this half an inch of copper scale, then push in another piece of gauze. Fasten the tube on a ring of the lamp stand as in fig. 5. Connect the tube to the supply pipe, and turn on a very small quantity of hydrogen. Light the gas at the other end and turn down the gas until the flame is only one-half inch high. Now slowly heat the copper scale with the Bunsen burner. Record all results and represent by equations.

2. Repeat exactly, using lead oxide in place of the copper oxide.

3. Water can be purified from suspended and insoluble substances by filtration. Dissolve some oxygen mixture in a test tube half full of water and filter into another test tube. What has been removed from the water? Evaporate the filtrate by

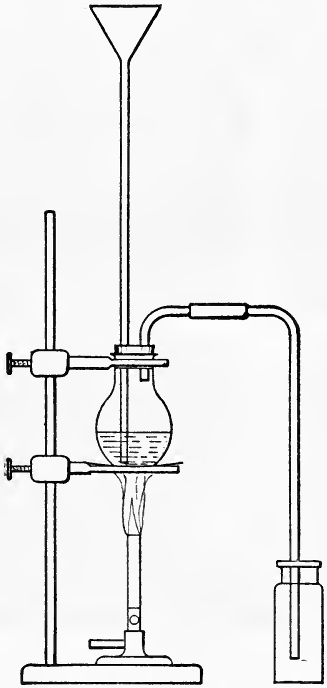


Fig. 6.

heating in the evaporating dish on the wire gauze. What is the residue?

4. Water can be purified from all kinds of matter by distillation. Fill the flask half full of water. Add a little sulphuric acid and some potassium permanganate solution. Arrange this on the wire gauze on the lamp stand with long glass delivery tube as in fig. 6. Heat to boiling and collect the distillate in a bottle. Taste the water that distills over.

5. Heat a small quantity of copper sulphate in a sealed tube. After the material is white, let some of the water run back on it.

6. Dissolve 1 s.s. of lead oxide in 10 c.c. boiling dilute hydrochloric acid. Pour the solution into water in another tube. This is a precipitate of lead chloride, soluble in hot water.

7. Repeat, using copper scale. The precipitate is cuprous chloride.

LABORATORY EXERCISE NO. 9.

Solution and Crystallization.

Materials. Sodium sulphate, $Na_2SO_4 + ?H_2O$; copper sulphate, $CuSO_4 + 5H_2O$.

Apparatus. Test tubes; funnel; filter paper; stirring rod.

Definitions—A **cold saturated solution** is one containing all of a given substance dissolved in it which it can absorb at that temperature. It can be made in two ways. Either by shaking an excess of the powdered substance with the liquid until no more will dissolve and filtering off the excess, or by cooling a hot solution and filtering off the substance that is deposited as a result of the cooling.

A **hot saturated solution** is one containing all of a given substance which it can dissolve at that temperature. It is best made by gradually adding the powdered substance to the boiling liquid until there is a decided excess present.

A **supersaturated solution.** When a hot saturated solution is allowed to cool, some of the material in solution generally crystallizes out. The amount that separates is the exact amount that the hot liquid dissolves more than the cold. Sometimes the excess refuses to crystallize when the hot saturated solution is cooled. The solution is then said to be **supersaturated.**

A supersaturated solution may be caused to crystallize either by dropping in a small crystal of the same material, or by scratching the inside of the vessel with a glass rod.

Mother liquor. The liquid that is left after the crystals that have separated from it have been removed.

Decant. To pour off a liquid gently so as not to disturb the sediment.

Use of filter flask. The filter flask is an arrangement for rapidly filtering off crystals or solid residues by means of suction.

Fit the rubber stopper with the porcelain sieve funnel tightly into the flask, connect the rubber tube to the suction pump and turn on the water. Test the apparatus by putting the palm of the hand firmly down over the funnel; a decided suction should be felt almost at once. Moisten a small disc of filter paper and fit it carefully over the sieve plate. Shake up the material to be filtered and decant it quickly into the funnel. It will be sucked dry almost at once. If necessary, and the substance admits of it, it may be washed free from the mother liquor by pouring successive small portions of water on it. The disc of filter paper with the substance is then removed, placed on a larger piece of filter paper and set aside to dry.

Preparation of a Supersaturated Solution of Sodium Sulphate.

1. Wash five test tubes and the funnel thoroughly. Prepare a filter and stand it in one of the test tubes.

2. Fill a dry test tube one third full of sodium sulphate. Heat a third of a test tube full of water nearly to boiling. Transfer the sodium sulphate to the hot water, cover with the thumb or with a rubber stopper and shake thoroughly for about three minutes.

3. Filter small portions into the several test tubes. Do not let any drops touch the side of the tube.

4. Cover each tube with a clean wet glass plate. Cool the tubes by standing them in water or by letting a stream of water run over them, without shaking.

5. If no crystals formed during the cooling, you have succeeded in getting a supersaturated solution of sodium sulphate. Into one tube drop a small crystal of the original salt. Scratch the inside of a second with a glass stirring rod. Allow some dust from the floor to fall into a third. Simply shake a fourth thoroughly.

6. The crystals that separate contain ten molecules of water of crystallization and are known as **Glauber's Salt.**

7. Filter off all of this material, making use of the filter flask. The salt is very soluble, therefore you cannot wash it. Set it aside to dry and explain next time what has happened to it.

8. Pulverize some copper sulphate and add it gradually to a t.t. half full of boiling water until, after boiling, a little of the salt remains undissolved. Filter this hot solution into several clean tubes. On cooling, the excess of salt will crystallize. If it does not, it can be made to do so by scratching the inside of the tube with a glass rod.

LABORATORY EXERCISE NO. 10.

The Action of Acids on Metals and on Basic Oxides.

64-a-b.

Materials. Iron wire; magnesium wire; zinc oxide; magnesium oxide.

Apparatus. Lamp stand; test tubes; funnel; filter paper; stirring rod.

Definitions—An **acid 55.** A **Base 56.** A **salt 58.** A **basic oxide 60.** **Water of crystallization,** see lab. ex. No. 8.

1. Fit a bulb-test tube with a rubber stopper and delivery tube. Bend up about 5 g. of iron wire in a bunch two inches long so it will go readily into the tube. Cover this with conc. hydrochloric acid. Fasten the tube in the clamp of the lamp stand at an angle of forty-five degrees so the bottom of the tube will come near the base of a Bunsen burner flame as in fig. 7. Place the flame near enough to the tube so the reaction will proceed readily without causing the liquid to boil over. Ignite the gas. Collect some of it over water, satisfy yourself that it is hydrogen. All iron contains sulphur and phosphorus as impurities and in this reaction these come out as the corresponding hydrogen compounds, H_2S and H_3P; both of these have very disagreeable odors. To get rid of them stand a second burner near the end of the delivery tube to keep the jet of gas burning.

2. When the action is nearly over, filter the liquid into a test tube. The filtrate should be green. Cool and crystallize, if possible. If it will not crystallize, boil it down until it does crystallize on cooling. The product is ferrous chloride, $FeCl_2 + 4H_2O$.

3. To three c.c. of dilute sulphuric acid in evap. dish add bits of magnesium wire until no more will dissolve. What gas was given off? Boil a little, cool, and drop in a small crystal of magnesium sulphate. The crystals contain seven molecules of water of crystallization and are known as **Epsom Salt.**

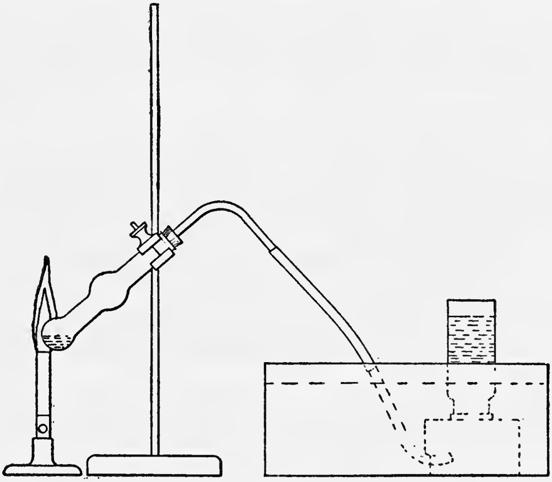


Fig. 7.

4. Make ferrous sulphate just as you made ferrous chloride, but in place of conc. acid use dil. sulphuric acid with an equal bulk of water. This product is known as green vitriol and has the formula $FeSO_4 + 7H_2O$.

5. Make magnesium chloride from the metal just as you made magnesium sulphate. The formula for the crystals is $MgCl_2 + 6H_2O$.

6. Write twenty-five equations showing the formation of the salts indicated in **64-a** by the above method.

7. Place about 10 c.c. of dilute sulphuric acid in beaker on wire gauze on lamp stand, add an equal amount of water, heat nearly to boiling and add zinc oxide little by little with much

stirring until no more will dissolve. Filter into t.t., add one or two drops of acid to filtrate, boil down and crystallize. Why was no hydrogen given off in this case?

8. Repeat, using magnesium oxide and sulphuric acid.

9. Write twenty-five equations, showing the formation of the same twenty-five salts by this method. **64-b.**

LABORATORY EXERCISE NO. 11.

The Action of Acids on Bases and on Carbonates.

64-c-d.

Materials. Sodium carbonate.

Apparatus. Lamp stand; wire gauze; evaporating dish; stirring rod; test tubes.

Definitions—An **alkali. 57.** A **base. 56.**

Neutral, reacting neither acid nor alkaline toward litmus.

Neutralization, the process of rendering a substance inactive toward litmus paper.

Preparation of a Salt by Neutralizing a Base with an Acid. **64-c.**

1. Place about 5 c.c. of sodium hydroxide solution from reagent bottle in porcelain dish on ring of lamp stand. **174.** Add an equal amount of water. Now add dilute hydrochloric acid little by little until a small piece of litmus paper just turns red. It is evident, now, that you have too much acid present. Take a small quantity of the sodium hydroxide solution in a test tube and dilute it with five times its bulk of water. Add this solution little by little with much stirring until the paper just turns blue. You now have too much alkali. Dilute some more of the hydrochloric acid with ten times its bulk of water and add it very cautiously until a fresh piece of litmus paper turns red. This solution will be neutral enough for our purpose, but if you are very careful it will be possible for you to show a red and a blue paper in the liquid at the same time; the liquid will then be quite neutral.

2. Filter the liquid, replace it in the evaporating dish and boil it down to dryness. Heat the dish red hot for five

minutes. Toward the end of the evaporation, if the salt sputters much, cover the dish with a piece of paper. The product is common salt. Taste it. If you were not particular to have the liquid either exactly neutral or very slightly acid, it will not taste like good salt.

3. Write twenty-five equations showing the formation of the same twenty-five salts by this method. **64-c.**

Preparation of Salts by Treating an Acid with a Salt of a More Volatile Acid. **64-d.**

One acid is said to be more volatile than another if, when the two are heated together, the first escapes before the second.

Sulphuric acid is one of the least volatile of acids. Therefore when it is heated with almost any salt, the other acid which that salt contains is set free.

Many acids are very unstable in the free condition, or cannot exist in the free condition at all, although they form perfectly stable and well known salts. When such salts are treated with sulphuric acid we get the decomposition products of the other acid in place of the acid itself.

These products are generally water and the acid anhydride. Carbonates are salts of carbonic acid. This is an unstable acid and cannot exist in the free condition. Therefore when a carbonate is treated with almost any acid we have carbon dioxide set free.

4. Dissolve ¼ t.t. full of sodium carbonate in one t.t. full of water in beaker, on wire gauze, on lamp stand with the aid of heat. Decant into two t.t.

5. To one portion, in beaker, on gauze, boiling hot, add dilute hydrochloric acid little by little, until effervescence nearly ceases on addition. Boil down as before.

Repeat using dilute sulphuric acid with other portion. Boil down and crystallize as with sodium sulphate in Ex. No. 9.

6. Write twenty-five equations showing the formation of the same twenty-five salts from the carbonates. **64-d.**

LABORATORY EXERCISE NO. 12.

The Formation of Salts by the Mixing of two Solutions that combine to form an insoluble Compound: also by fusing together Basic with Acidic Oxides. **64-e-f.**

Materials. Barium chloride sol.; calcium chloride; sodium carbonate; sand; litharge; potassium chromate; lead nitrate.

Apparatus. Test tubes; sealed tubes; filter flask.

Solubility rule. All the hydrogen, sodium, potassium and ammonium compounds are soluble, also the chlorates, acetates and nitrates and all the chlorides except those of silver, lead and mercurous mercury. **165, 166.**

1. Dissolve separately about five grams each of sodium carbonate and calcium chloride. Heat the solutions boiling hot and mix them in a bottle. Filter with the suction pump, using a clean filter flask. The filtrate should contain a slight excess of sodium carbonate. If the filtrate is not perfectly clear, filter through an ordinary filter. Acidify the filtrate with hydrochloric acid and boil it down to dryness in evaporating dish.

When mixing reagents in a test tube always add the second reagent a drop at a time until you ascertain what the result is going to be, then add more and mix thoroughly. It is well also to use only a small quantity of the first reagent and dilute it with water if a larger bulk is desired.

Make the following mixtures in test tubes:

2. Barium chloride with sulphuric acid; potassium chromate with barium chloride; potassium chromate with lead nitrate; lead nitrate with sulphuric acid; sodium carbonate solution with barium chloride.

3. Write equations showing the formation of the same twenty-five compounds by this method, making use of silver chloride, barium sulphate or calcium carbonate as the insoluble compound. **64-e.**

4. Grind a very small quantity of sand in mortar to an impalpable powder, mix it thoroughly with an equal bulk of lead oxide and heat in a sealed tube. The sand is an acidic oxide, which at a high temperature unites readily with the basic

lead oxide, forming a lead silicate. This is a variety of glass and is absorbed into the glass of the sealed tube.

5. Write equations showing the formation of the silicates, SiO_3'', chromates, CrO_4'', sulphates, and phosphates of the same five positive radicals, by this method. **64-f.**

LABORATORY EXERCISE NO. 13.

The Preparation and Properties of Chlorine.

Materials. Manganese dioxide; sodium chloride; potassium chlorate; chlorine water; red cloth; paper with printing; paper with writing ink marks; silver nitrate.

Apparatus. Test tubes.

Caution. When smelling unknown substances, first fill the lungs with air and then waft some of the suspected odor toward the nose. Do not smell the pure gas.

In this exercise, as well as all others in which disagreeable odors are produced, the hoods should be used as much as possible.

Preparation. Chlorine is prepared in the laboratory by removing the hydrogen from hydrochloric acid. That is by oxidizing hydrochloric acid, and it may be accomplished by almost any oxidizing agent. **74-76-78.**

1. To ½ s.s. of powdered manganese dioxide in t.t. add ½ inch of dil. hydrochloric acid, shake and warm gently. Note color, odor and action on moist litmus of the escaping gas. **75.**

2. To ½ inch of conc. hydrochloric acid in t.t. add a few drops of conc. nitric acid. Warm and note what gas is given off. **77.**

3. To ½ inch of conc. hydrochloric acid in t.t. add ½ s.s. of potassium chlorate. Warm and note the result. Is the gas the same in all three of these cases?

4. Place in t.t. 2 s.s. of common salt and 2 of manganese dioxide. Mix them thoroughly by shaking and add four drops of water; shake again. Then add a little conc. sulphuric acid. What gas is given off? Is it the same as the gas that comes from the conc. hydrochloric acid bottle?

5. Place in t.t. one l.s. of sodium chloride, four drops of

water and about three c.c. of conc. sulphuric acid. Note large volumes of a colorless gas that fumes in air.

6. Assuming this gas to be hydrochloric acid, what ought to be the effect of mixing manganese dioxide with the salt before adding the sulphuric acid? Explain the reaction.

7. Place in chlorine water in t.t. some red cloth, paper with printing and paper with writing ink marks. Chlorine bleaches because it oxidizes in the presence of water. **71.**

8. To a few drops of silver nitrate solution add some chlorine water, to another portion add some hydrochloric acid and to a third add some sodium chloride solution.

9. Shake the tubes thoroughly to cause the precipitate to settle and decant the excess of liquid. Add ammonium hydroxide to the residue. Then add dilute nitric acid carefully to the ammonium hydroxide solution.

10. Write equations showing the formation of chlorine from hydrochloric acid and ten different oxidizing agents. **75-76-78.** Also write five equations showing the preparation of chlorine from a chloride and five different oxidizing agents in the presence of sulphuric acid.

LABORATORY EXERCISE NO. 14.

Preparation and Properties of Hydrochloric Acid. **64-d.**

Materials. Sodium chloride; iron filings; silver nitrate sol.

Apparatus. Lamp stand; flask; wire gauze; funnel tube; glass bends; bottles; rubber connectors; rubber stoppers.

1. Set up the apparatus as indicated in the sketch, fig. 8. A contains about 20 g. of sodium chloride; B is empty; C contains water, but the surface must not quite touch the glass tube; D has no rubber stopper and contains 10 c.c. of ammonium hydroxide solution. The funnel tube must go to the bottom of the flask.

2. Add 8 c.c. of water to the flask through the funnel and be sure that it wets the salt all through. Then add 25 c.c. conc. sulphuric acid.

3. B becomes filled with HCl gas at once. The gas is absorbed in the water in C. What are the white fumes in D?

4. Place a low flame under the flask, remove the bottle B, and replace it with another. Fill several bottles in this way with HCl gas. Keep them right side up and covered with glass plates. In the meantime notice what is going on in the water just below the end of the tube in C.

5. Draw out the end of a glass tube eight or nine inches long to a very small size but be sure there is a hole through it. Pass the drawn out end through one hole of a two-hole stopper and

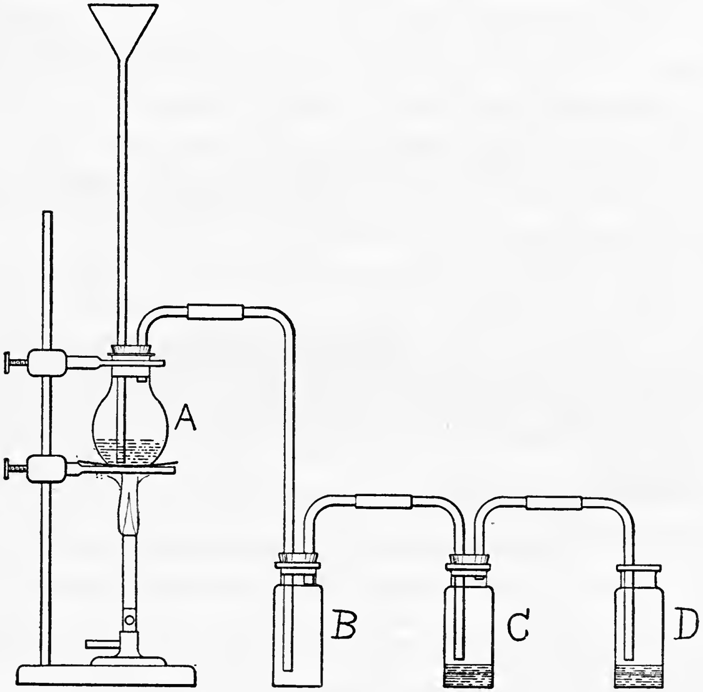


Fig. 8.

stop up the other hole with a sealed tube; quickly insert the stopper in one of the bottles filled with HCl gas and invert it so the long end of the tube will project downward into a large bottle filled with water. The water will soon rise and produce a fountain in the bottle. If it is slow to start, allow some cold water to run over the acid bottle. See fig. 9.

6. Place about 5 c.c. of strong ammonium hydroxide in a small bottle, shake it about and pour it out. Invert over it a

covered bottle filled with HCl gas. Withdraw the glass plate between the two.

7. Test a bottle of the gas with a burning splint to see if it supports combustion.

8. Test the liquid in C with litmus paper. Note the odor of it. Does it smell like the conc. hydrochloric acid in reagent bottle? Pour some of it on iron filings in t.t. and prove that the gas given off is hydrogen. Test a portion of it with silver nitrate and note that the precipitate dissolves in ammonium hydroxide. Heat some with a little manganese dioxide and note that chlorine is given off.

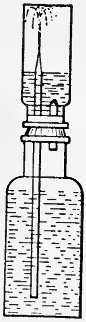


Fig. 9.

9. The white solid which separates in the flask is acid sodium sulphate, $HNaSO_4$. See Acid salt, Exp. 25.

10. Write equations showing the formation of hydrochloric acid from ten different chlorides. **64-d.**

LABORATORY EXERCISE NO. 15.

The Preparation and Properties of Bromine and Iodine.

75-76.

Materials. Potassium bromide; potassium iodide; starch; manganese dioxide; carbon disulphide; bromine water; chlorine water; alcohol; red calico; iodine.

Apparatus. Test tubes; cork stopper with a hole in it; medium glass bend.

Definitions. Sublimation. The process of distilling a solid in such a way that the vapor condenses directly to the solid form.

Sublime. To distill a solid in such a way that the vapor condenses directly to the solid form.

Sublimate. The solid product formed by sublimation.

1. Place in a bulb-test tube 2 s.s. of potassium bromide, two of manganese dioxide and ten drops of water.

2. Fit the tube with **cork** stopper and medium glass bend as in fig. 10. Fill a large bottle with water and stand in it a t.t. containing a few c.c. of water. Place the end of the glass bend well down in the test tube and see that it does not quite touch the water.

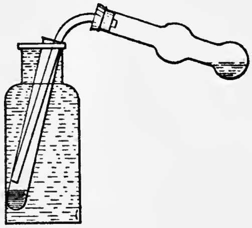


Fig. 10.

3. Remove the cork stopper and add about 5 c.c. of conc. sulphuric acid to the mixture of manganese dioxide and potassium bromide. Replace the stopper firmly. The whole will remain supported by itself.

4. Holding the burner in the hand, warm the tube throughout its whole length until five or six drops of bromine have distilled over. **Do not inhale the vapor of bromine. Do not get liquid bromine on the hands.**

5. Remove the tube from the bottle and examine the drop of bromine. Fill the tube half full of water and shake thoroughly to dissolve the bromine and form bromine water. Use this bromine water in the following experiments. If you need more bromine, use the bromine water from the reagent bottle.

Caution. Carbon disulphide is extremely inflammable. Do not have a flame near when using it in the next experiment.

6. Test the bleaching power of bromine water just as you did with chlorine water. Add some carbon disulphide, about

4 c.c. to a tube one quarter full of water and shake it thoroughly. Note that the carbon disulphide does not dissolve in the water, but ultimately settles out in a clear globule at the bottom of the tube. Now add some bromine water and shake again. Bromine is soluble in carbon disulphide, and since the latter is insoluble in water, the color of the bromine is soon transferred to it, changing it to a yellow or a red, according to the quantity of bromine present. This is a common test for free bromine.

7. Dissolve 1 s.s. of potassium bromide in about 5 c.c. of water in a test tube, add carbon disulphide and shake. There is no color change, therefore no free bromine. Now add some chlorine water and shake again.

8. Fit a bulb-test tube with a cork stopper and straight glass delivery tube three inches long. Place in the tube 1 s.s. of

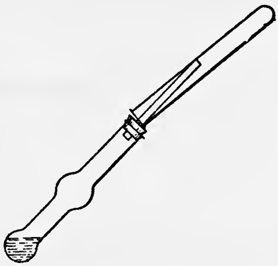


Fig. 11.

potassium iodide, two of manganese dioxide and ten drops of water. Add about 5 c.c. of conc. sulphuric acid. Replace the stopper firmly and invert a dry test tube over the delivery tube. See fig. 11.

9. Holding the generating tube in the wire test tube holder, heat it carefully at first and finally boil the liquid to drive the iodine up into the inverted tube. The action is over when the purple color disappears from the generating tube.

10. Fill the receiving tube half full of water, cover with the thumb and shake thoroughly. What does this show you about the solubility of iodine in water? The brown solution is iodine water and may be used in some of the following experiments.

11. Place a small crystal of iodine (from the dish) in a dry t.t. and heat it gently. Does it melt, boil, vaporize, or sublime? Let the tube cool and examine the inside carefully. Fill it half full of water and heat it to boiling. Is iodine volatile with steam? Add some carbon disulphide and iodine water to a tube half full of water and shake thoroughly. This is a common test for free iodine.

12. Dissolve ½ of a s.s. of potassium iodide in 5 c.c. of water in test tube. Add carbon disulphide and a little chlorine water and shake thoroughly. What did the chlorine do to the potassium iodide? Repeat, using bromine water.

13. Heat a t.t. full of water in beaker to boiling. Mix a s.s. of starch with 5 c.c. of water in test tube and pour the mixture into the boiling water. This forms a starch paste solution. To a very small quantity of this solution in a test tube half full of water add a few drops of iodine water. This is the best test for free iodine.

14. Test the solubility of iodine in alcohol, also in water containing a little potassium iodide. A solution of iodine in alcohol is known as **tincture** of iodine.

15. Pulverize ½ of a s.s. of iodine, place it in test tube one-quarter full of water and pass in hydrogen sulphide until all the iodine has dissolved. What is the residue? **71.** Filter it off and test the filtrate for an acid. It contains hydriodic acid.

16. Write equations showing the formation of bromine and iodine from five different bromides and five different iodides by the use of sulphuric acid and manganese dioxide.

LABORATORY EXERCISE NO. 16.

Sulphur.

Materials. Sulphur; carbon disulphide; iron filings.

Apparatus. Test tubes; evaporating dish; combustion spoon.

Note. Sulphur is insoluble in water. Do not attempt to wash with water, tubes that have contained melted sulphur.

1. Grind together in the mortar about a gram of sulphur and 5 c.c. of carbon disulphide. Filter a few drops of this through a dry folded filter paper held in the hand upon a glass plate. After the liquid has evaporated examine the crystals under the microscope.

2. Spread a large piece of paper on the table. Stand the burner on this in an inclined position. Fill a dry test tube two-thirds full of clean lumps of sulphur. Holding it in the test tube holder, heat it **very slowly and carefully** until the sulphur is all melted to a light straw color. If it becomes dark colored you have heated it too fast. Pour some of this light-colored liquid into water. Now heat it very much hotter and note the changes that it undergoes. When it begins to boil, pour it rapidly into a large quantity of water. Remove the product from the water and examine it at once. Keep this product for several days and note the changes.

3. Let a tube half full of melted sulphur stand until it begins to solidify and crystals are seen lining the inside of the tube. Then pour out the liquid sulphur from the center of the tube. **Don't pour the liquid sulphur into the sink.**

4. Place a small quantity of sulphur in the bowl of the combustion spoon, not enough to fill it when melted, and heat the spoon until the sulphur is burning violently. Lower the burning sulphur into a bottle. Keep the bottle covered with a glass plate. Let the sulphur burn as long as it will. What product is formed here? **69.** Is it a solid, a liquid or a gas? Is it acidic or basic? Will it dissolve in water? What effect has it on litmus paper?

5. Weigh out approximately 6 g. of iron filings and 3 g. of sulphur. Grind these together thoroughly in the mortar. Transfer the mixture to the tube in which you melted the sulphur and from which as much sulphur has been removed as possible by pouring it out in the melted condition. Heat the bottom of the tube as hot as possible in the hood until the reaction begins and then remove it from the flame and observe the progress of the reaction. The compound that is formed is ferrous sulphide. When the tube is cool, break it and remove the solid lump. Save it for the next experiment.

6. Write ten equations showing the formation of sulphides by the direct union of sulphur with metals. **40.**

LABORATORY EXERCISE NO. 17.

Preparation and Properties of Hydrogen Sulphide and Sulphur Dioxide.

Materials. Ferrous sulphide, from last experiment. Cupric sulphate sol.; barium chloride sol.; lead nitrate sol.; potassium permanganate sol.; red cloth; potassium bichromate sol.; cadmium sulphate; cylinder of liquid sulphur dioxide.

Apparatus. Bottles; beaker; evaporating dish; test tubes; stopper; delivery tube.

As far as possible all experiments with hydrogen sulphide and with sulphur dioxide must be performed in the hoods.

1. Treat the lump of ferrous sulphide obtained in the last experiment in a t.t. with dilute sulphuric acid and warm to hasten the reaction. Fit the tube with a rubber stopper and delivery tube and collect some of the gas in a test tube by displacement of air. Ignite the gas. What are the products? **68.** Represent by equation.

2. Pass some hydrogen sulphide from the supply bottle in the hood into a test tube half full of water containing a little cupric sulphate solution. Filter off the precipitate and pass the gas again into the filtrate. Ascertain whether all the copper can be removed in this way.

3. Repeat, using lead nitrate solution. Also cadmium sulphate solution, formula, $CdSO_4$.

4. Test the action of a water solution of hydrogen sulphide on litmus paper. What would be formed if hydrogen sulphide were passed into a solution of sodium hydroxide? Ammonium hydroxide?

5. Fill several bottles with sulphur dioxide from the cylinder in the hood. Keep them right side up and covered with glass plates. Be sure to turn off the gas at the cylinder.

Invert one of the bottles in a dish of water. Fill another bottle one-quarter full of water, cover tightly with the palm of the hand and shake thoroughly; note the suction on the hand. What does this indicate? Test the liquid with litmus

paper. What acid is in the water? Represent its formation by an equation. Save the solution.

6. Add a few drops of conc. nitric acid to a bottle containing sulphur dioxide. **73.** Pour some potassium permanganate solution into some sulphur dioxide water. Also some potassium bichromate solution. These are both oxidizing agents and change the sulphur dioxide to sulphur trioxide, which remains dissolved in the water in the form of sulphuric acid. Indicate these reactions by equations. **75-76.**

7. Neutralize some sulphur dioxide water with sodium hydroxide solution and evaporate it to dryness in the porcelain dish. **64-g.** The product is sodium sulphite, Na_2SO_3. If this be treated with an acid, sulphur dioxide will be given off.

8. Represent the formation of hydrogen sulphide from five different sulphides by the action of acids.

Represent the formation of sulphur dioxide from sulphuric acid by the action of five different metals. **71.**

Represent the oxidation of hydrogen sulphide to sulphuric acid by five different oxidizing agents. **78.**

Represent the oxidation of sulphur dioxide to sulphuric acid by five different oxidizing agents.

LABORATORY EXERCISE NO. 18.

Preparation and Properties of Sulphuric Acid.

Materials. Sulphur; barium chloride sol.; bromine water; cupric sulphate sol.; zinc sulphate; iron wire; potassium permanganate sol.

Apparatus. Combustion spoon; test tubes; bottles.

Caution. Do not spill sulphuric acid.
Do not pour water into sulphuric acid.
Do not pour boiling sulphuric acid into water.

1. Place about ½ inch of water in a large bottle. Place a small quantity of sulphur in the combustion spoon, less than will fill it when melted, and heat it until it is all melted and blazing. Suspend the blazing sulphur directly over the water, keeping the bottle covered, until it will burn no longer. Remove

the spoon, cover the bottle with the hand and shake thoroughly to dissolve the sulphur dioxide in the water. This forms a solution of sulphurous acid. To a portion of this solution add a few drops of barium chloride solution. What does the slight white precipitate indicate? Then add a few drops of conc. nitric acid. What does the additional white precipitate indicate? **73.**

2. Repeat, using bromine water in place of nitric acid. **71.**

3. Repeat, using potassium permanganate sol. **76.** Write all equations.

4. Notice the weight of the conc. sulphuric acid in the reagent bottle, also its oily nature. Pour some conc. sulphuric acid into 10 c.c. of water in a test tube. Put pieces of wood and paper in hot conc. sulphuric acid. Dilute some sulphuric acid with ten volumes of water and write on paper with it. Heat the paper.

5. Ascertain the effect of cold and hot conc. sulphuric acid on iron wire.

Hot conc. sulphuric acid is an oxidizing agent according to the equation, $H_2SO_4 = H_2O + SO_2 + O.$

$$O + Fe = FeO.$$

$$FeO + H_2SO_4 = FeSO_4 + H_2O.$$

$$2H_2SO_4 + Fe = FeSO_4 + SO_2 + 2H_2O.$$

What is the effect of hot dilute sulphuric acid on iron wire?

6. Systematically dilute some of the acid in ten test tubes. Find approximately how dilute a solution will give a test with barium chloride.

7. Test solutions of cupric sulphate and zinc sulphate with barium chloride.

8. **Test for sulphuric acid or a soluble sulphate.** Add barium chloride and some hydrochloric acid to the solution to be tested; a white precipitate insoluble in hydrochloric acid is evidence of the presence of the sulphate radical.

Neutral and alkaline solutions of phosphates give a white precipitate of barium phosphate, but this is soluble in hydrochloric acid, hence the use of acid in the above test.

9. Write equations showing the oxidation of sulphur, hydrogen sulphide and ammonium sulphide to sulphuric acid by means of nitric acid. **73.**

LABORATORY EXERCISE NO. 19.

The Preparation and Properties of Nitrogen and Ammonia.

Materials. Yellow phosphorus in beads **176**; sodium nitrite; ammonium chloride; calcium oxide; ammonium nitrate; ammonium sulphate.

Apparatus. Water pan; 1000 c.c. bottle; rubber stopper and delivery tube; flat cork.

1. Place a small piece of yellow phosphorus in the hollow of a flat cork, float on water in water pan over shelf, ignite the

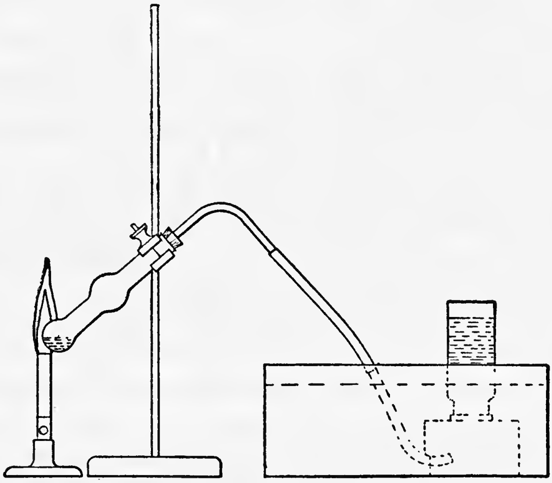


Fig. 12.

phosphorus, invert large bottle over it at once and let it stand until all fumes have disappeared. Why did bubbles of air come out of bottle? Why does water enter bottle? What proportion of bottle becomes filled with water?

The white smoke is phosphorus pentoxide, P_2O_5.

The gas left in the bottle is atmospheric nitrogen. Note the lack of color, odor and the fact that it will not support combustion.

2. Chemical nitrogen may be prepared by heating a solution of ammonium nitrite. As this is not ordinarily kept in stock,

a mixture of sodium nitrite and ammonium chloride is used. Fit a test tube with a rubber stopper and delivery tube. See fig. 12. Place in it 1 s.s. of ammonium chloride and an equal amount of sodium nitrite with 5 c.c. of water. Heat the mixture carefully and collect the gas over water. Indicate the reaction by two equations.

3. To 1 s.s. of ammonium chloride in t.t. add a small quantity of sodium hydroxide sol. from reagent bottle, and warm gently. Note the odor and effect of the gas on moist litmus paper; on a rod wet with conc. hydrochloric acid. **Do not dip the rod into the reagent bottle.**

4. Repeat, using ammonium nitrate; and again using ammonium sulphate. The gas set free is ammonia, NH_3.

5. When an ammonium compound is heated it breaks up into ammonia and the corresponding acid. **137.** As the gases cool they reunite to form the original solid substance.

Heat a gram or two of ammonium chloride in a dry test tube and insert in the mouth of the tube a piece of moist red litmus paper. Explain the color change on the paper. **99.** What is the reason for using an alkali or an alkaline oxide with an ammonium compound to form ammonia?

6. Mix thoroughly in the mortar about 5 g. of ammonium chloride and 10 g. of calcium oxide. Transfer the mixture to a test tube, adding only enough to half fill it. Fit the tube with a rubber stopper and medium bend. Tap the tube in a horizontal position to obtain an air space over the whole length of the mixture. Fix it in the clamp on the lamp stand in a nearly horizontal position with the stopper end lower, as in fig. 13, so that any water formed will not run back on the hot glass. Heat very carefully, holding the burner in the hand, **but not hot enough to cause the flame to turn yellow, as a result of melting the tube.** Collect the ammonia in **dry** bottles by upward displacement of air.

7. Make an ammonia fountain exactly as you made the hydrochloric acid fountain in Ex. 14, 5.

8. Invert a bottle or tube of ammonia in a dish of water. Drop some conc. hydrochloric acid into a bottle of ammonia. Test ammonia gas with moist litmus.

9. Neutralize some ammonium hydroxide from the reagent bottle with nitric acid. Boil down the solution in the evapo-

rating dish until some of it will solidify on the end of a glass rod. What have you made? Save the product.

10. Write equations showing the formation of ammonia from ten different ammonium compounds, using at least four different alkalies.

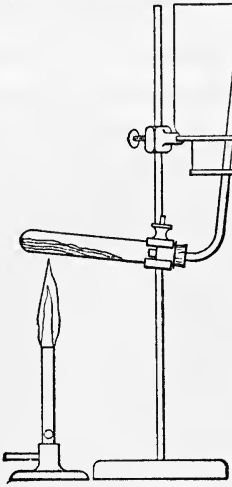


Fig. 13.

11. Show by equations the formation of ten different ammonium compounds by the direct union of ammonia with ten different acids.

LABORATORY EXERCISE NO. 20.

Some Oxides of Nitrogen and Nitric Acid.

Materials. Ammonium nitrate; scrap copper; sodium nitrite; phosphorus; sodium nitrate.

Apparatus. Lamp stand; water pan; bottles; rubber stopper and delivery tube; funnel tube.

1. Place in test tube about 1½ inch of ammonium nitrate. Fit tube with stopper and delivery tube. Fix tube in clamp on lamp stand in an inclined position so the bottom of tube

comes near the base of the Bunsen flame. See fig. 12. Heat very carefully until the material is all melted, then only enough to produce a steady evolution of the gas. Excessive heating may produce an explosion. Under no circumstances should the material be heated until it is nearly all gone. The gas given off is nitrous oxide, N_2O. Remove delivery tube from water pan before removing the flame.

2. Test the solubility of this gas in water. Make the same experiments with it that you did with oxygen. How could you distinguish this gas from oxygen?

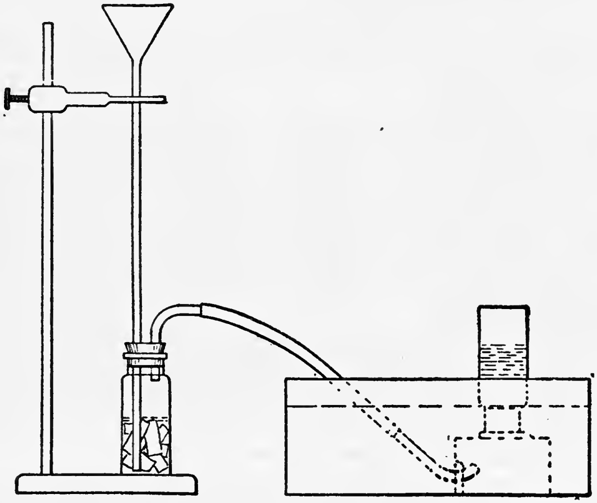


Fig. 14.

3. Fit a small bottle with a two-hole rubber stopper. Through one of the holes pass a funnel tube. The end of the tube must reach very near to the bottom of the bottle. Through the other pass a short bend, to which a delivery tube is attached. Place several pieces of scrap copper in the bottle and stand the whole on the lamp stand with the funnel tube passing through a ring to prevent the apparatus from being upset. See fig. 14. Arrange to collect the gas over water in the water pan. Add about 20 c.c. of dilute nitric acid and then conc. nitric acid little by little as necessary to produce the reaction. Have all receiving bottles upside down in the pan, and absolutely full of

water. Do not allow any air to enter when removing the bottles of gas from the pan.

What is the color of this gas? It is called nitric oxide, NO.

Open a bottle of it in the air. Why was the gas colored brown in the generating bottle at first? Will it support combustion? Test it with very hotly burning phosphorus in the hood. What would be the effect of mixing nitric oxide with oxygen? Nitric oxide with nitrous oxide?

4. Write ten equations showing the action of nitric acid on ten different metals when nitric oxide is formed.

5. Treat sodium nitrite with several different acids in test tube. Nitrous acid, HNO_2, is unstable and cannot exist in

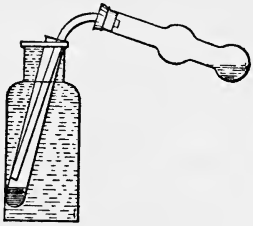


Fig. 15.

the free condition. Therefore when it is formed in any reaction we get its decomposition products.

$$2HNO_2 = H_2O + N_2O_3.$$

6. Fill a large bottle with water. Place in it a test tube. Fit a bulb-test tube with a **cork** stopper through which passes the short arm of a medium glass bend. Place in this tube 3 s.s. of sodium nitrate, five drops of water and about 5 c.c. of conc. sulphuric acid. Replace the stopper firmly. See fig. 15.

Place the long end of medium glass bend in receiving tube. The whole apparatus will support itself. Holding the burner in the hand, heat the mixture carefully until about 1 c.c. of nitric acid has distilled over. Satisfy yourself that this liquid is nitric acid by its characteristic action on copper wire.

7. Write ten different equations showing the preparation of nitric acid by the action of sulphuric acid on nitrates.

LABORATORY EXERCISE NO. 21.

Examination of an Unknown Substance.

The object of this exercise is to impress on the student again, the necessity for close original observation in all his chemical work. To call his attention again to the fact that every simple experiment that he has done or is going to do is simply a practical way of answering some question concerning the material he is operating on, and it is for him to see what the question is in every case.

And that his notes are simply a series of definite *facts* which he has learned by his own observation of the experiments done in answer to the various questions he has had in mind.

Materials. Several unmarked chemical compounds; Silver nitrate sol.; barium chloride sol.; potassium chromate sol.

Apparatus. Sealed tubes; test tubes.

Definitions—Sublime. A substance is said to sublime when it changes from the solid to the gaseous form and back to the solid on being distilled. **Water of crystallization**, see lab. ex. No. 8. **Precipitate.** Ex. No. 8.

Never attempt to write equations unless there is some definite reaction, the nature of which is known.

1. Will it melt, boil, vaporize or sublime? Place a small quantity in a sealed tube and heat it as hot as possible, noting results. Don't mistake a transparent melted solid for water. If any water is present it will distill up on the sides of the tube or perhaps come out in the form of steam.

Leave at least one blank line after every recorded observation where there was a definite reaction, to be filled in later, when the nature of the substance is known, with the equation representing that reaction.

2. Is it soluble in water, and if so to what extent? Use very small quantities in testing solubility. Make the test in both cold and boiling water. Is it soluble in dil. nitric acid? How does the water solution affect litmus paper?

3. How does it affect the flame? Heat some on an iron wire loop which has been thoroughly cleaned by dipping it in hydrochloric acid and washing it. **Don't clean iron wire by dipping it in the reagent bottles.**

4. Treat portions of the dry substance, in test tubes, with conc. sulphuric acid, dil. hydrochloric acid, and sodium hydroxide sol., warming in each case.

5. Test portions of the water solution or dilute nitric acid sol. with silver nitrate, barium chloride, potassium chromate, hydrogen sulphide gas, and sulphuric acid.

When one unknown substance is finished exchange it for another.

Work very slowly and be sure of your facts before recording them. Repeat tests several times if necessary.

LABORATORY EXERCISE NO. 22.

Phosphorus and Arsenic.

Materials. Yellow phosphorus in beads **176**; carbon disulphide; red phosphorus; hydrogen disodium phosphate sol.; silver nitrate sol.; arsenic trioxide; charcoal; metallic arsenic.

Apparatus. Iron plate; lamp stand; large dry bottle; test tubes.

1. Dry a piece of yellow phosphorus by touching it to filter paper. Place it on a dry iron plate on ring of lamp stand. Ignite it and cover it at once with a large **dry** bottle. Let it stand until it ceases to burn and the white product has settled. Scrape out the phosphorus pentoxide with a wire, place it on a glass plate and allow one drop of water to fall on it. Breathe on some of the product. Do you consider that it has a great affinity for water? Test it with moist litmus paper. What acids does phosphorus form? Which one was formed by the union of the pentoxide with the one drop of water?

2. Dissolve one piece of yellow phosphorus in about 4 c.c. of carbon disulphide in t.t. Pour the solution on disc of filter paper supported on ring of lamp stand in hood and wait developments. **Don't spill this solution on your clothes or on your hands.**

3. Dissolve ¼ of a s.s. of red phosphorus in about 5 c.c. of conc. nitric acid in t.t. with the aid of heat. The nitric acid is a strong oxidizing agent and oxidizes the phosphorus to the pentoxide, which dissolves in the water present to form phos-

phoric acid, H_3PO_4. **73.** Dilute the solution with water, add a few drops of silver nitrate, mix these two, then add ammonia carefully so as not to mix. There ought to be a yellow ring of silver phosphate in the neutral zone. What does this show about the solubility of silver phosphate in both ammonia and nitric acid? To appreciate better the nature of silver phosphate, add some silver nitrate to a solution of hydrogen disodium phosphate.

4. Warm a little piece of metallic arsenic in a sealed tube. Let the tube cool. Shake out the arsenic and transfer it to another sealed tube. Heat as hot as possible until no further change takes place. Note the two modifications of arsenic.

5. Heat a **very minute** piece of arsenic very carefully and slowly in a dry test tube. Explain what takes place. Can any crystals be seen in the tube?

6. Heat a very small quantity of arsenic trioxide in a sealed tube. Place ½ of a s.s. of arsenic trioxide in a sealed tube and above it about 1½ inch of broken charcoal. Heat the charcoal red hot first, and then the arsenic.

7. Test the solubility of arsenic trioxide in acids and in alkalies. **63.** Is it basic or acidic? Pass hydrogen sulphide into the hydrochloric acid solution.

8. Write equations showing the oxidation of phosphorus and arsenic to the corresponding acids, each by five different oxidizing agents. **73-75.**

LABORATORY EXERCISE NO. 23.

Carbon.

Materials. Powdered charcoal; animal charcoal; litmus solution; lime water; lead oxide; soft coal; copper scale; solution of sulphate of quinine; potassium dichromate, $K_2Cr_2O_7$.

Apparatus. Lamp stand; iron crucible; test tubes; sealed tubes; beaker.

1. Heat bits of wood and paper in sealed tubes and note that the volatile gases burn. Also that something like water is given off that has an acid reaction toward litmus paper. What is left in the tube in each case?

2. Fill iron crucible half full of animal charcoal, cover with iron plate, and heat it red hot for five minutes. While this is being heated make a solution of hydrogen sulphide by passing the gas into water in a test tube. Add 5 c.c. of the charcoal to 5 c.c. of the hydrogen sulphide solution. Shake thoroughly and filter. Note that the odor has been entirely removed from the water.

3. Boil 5 c.c. of the freshly ignited charcoal with 5 c.c. of litmus solution for two minutes and filter.

4. Shake a solution of quinine with some of the charcoal, filter and taste the filtrate.

5. Shake some potassium bichromate solution with charcoal and filter through a dry filter.

6. What classes of bodies does charcoal remove from solution?

7. Place in a sealed tube an intimate mixture of lead oxide and powdered wood charcoal, connect the tube to a medium glass bend, heat as hot as possible and pass the evolved gas into lime water. Is there any evidence of lead in the tube? **80.**

8. Repeat with copper scale.

9. Place some soft coal in a sealed tube and heat as hot as possible. What is the product left in the tube?

10. Write ten equations showing the reduction of ten common metals from their oxides. **82.**

LABORATORY EXERCISE NO. 24.

The Preparation and Properties of Carbon Dioxide.

Materials. Broken marble; charcoal in lumps; wood; oxygen; magnesium ribbon; lime water; sodium carbonate; sodium bicarbonate; carbon dioxide in cylinders.

Apparatus. Lamp stand; bottles; funnel tube; wire gauze; evaporating dish; combustion spoon.

1. Fill a bottle with oxygen by displacement of air. Test with splint to see when it is full. Add a little lime water. Ignite a piece of charcoal on the combustion spoon and lower it into the gas. What is the white precipitate?

2. Place several pieces of broken marble in small bottle, fit it with two-holed stopper, funnel tube and delivery tube. Stand it on the lamp stand with the funnel tube passing through one

of the rings to prevent it being upset. Connect it to a wash bottle as indicated in fig. 16. Add dilute hydrochloric acid through the funnel tube and collect several bottles of the gas by displacement of air. Test with a burning splint to see when the bottles are full.

3. Holding a piece of magnesium wire in the forceps, ignite it and thrust it into a bottle that is partly filled with the gas. What are the black specks?

4. Pour a few drops of lime water into a bottle containing

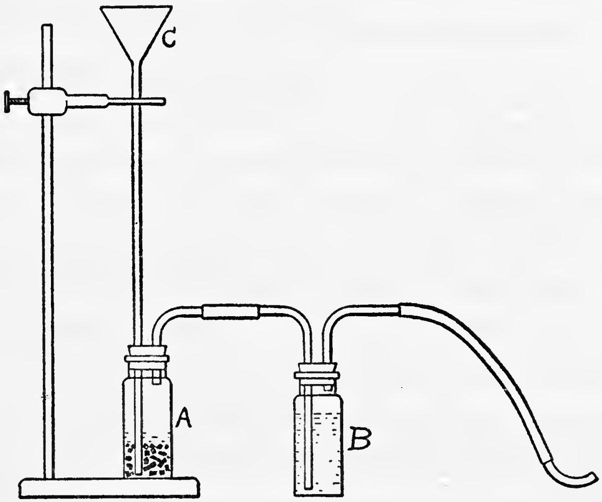


Fig. 16.

some of the gas. Blow through lime water in wash bottle. What does this show about carbon dioxide in the breath?

5. Arrange the wash bottle so that the suction pump will suck air through lime water. Allow the air to pass for five minutes. What does this show about the amount of carbon dioxide in the air compared to that in the breath?

6. Allow a slow stream of the gas to bubble through 10 c.c. of lime water in a test tube. Note and explain all changes. Boil the final clear solution.

7. Pass a stream of the gas into a very small quantity of sodium hydroxide in a test tube. What product is formed? **64-g.**

8. Treat sodium carbonate with acids and represent the reactions by equations. **64-d.**

9. Heat some bicarbonate of soda in a test tube or a sealed tube and pass the gas into lime water.

10. Decant the liquid from the marble in the generator, wash the marble and return it to the dish. Boil down the liquid or a portion of it in the evaporating dish over the wire gauze until the residue is perfectly dry. What is it? Will it dissolve in water?

11. Write equations showing the decomposition of ten different carbonates by ten different acids. **64-d.**

LABORATORY EXERCISE NO. 25.

Acid Salts and Baking Powder.

Materials. Carbon dioxide; acid potassium tartrate; bicarbonate of soda; corn starch; tartaric acid; normal potassium tartrate; lime water.

Apparatus. Test tubes; glass bend; rubber connector; bottles; beaker; large porcelain mortar.

Definitions—A **polybasic acid** is one having more than one replaceable hydrogen. Ex. H_3PO_4; H_2SO_4.

A **diabasic acid** is one having two replaceable hydrogens. Ex. H_2SO_4.

A **tribasic acid** is one having three replaceable hydrogens. Ex. H_3PO_4.

An **acid salt** is one formed by replacing part of the hydrogen of a polybasic acid by a basic radical. $HNaSO_4$; HNa_2PO_4.

A **normal salt** is formed by replacing all the hydrogens of a polybasic acid by the same kind of positive radical. Ex. Na_2SO_4.

A **basic salt** may be considered as one formed by replacing only part of the oxygen of a basic oxide or part of the hydroxide of a base by an acid radical.

1. Stand a test tube one-half full of sodium hydroxide solution in a small bottle of water. By means of a medium glass bend let a steady stream of carbon dioxide bubble through it until a finely divided heavy white precipitate has formed. **64-g.**

This will take nearly forty minutes. Filter it off by means of suction, using the filter flask and pump, set it aside to dry and

satisfy yourself at the next laboratory exercise that it is bicarbonate of soda. Heat some in a sealed tube and pass the gas into lime water.

2. Weigh out 10 g. of normal potassium tartrate, $K_2C_4H_4O_6$. From the following equation, figure out the proper quantity of tartaric acid, and weigh it out:

$$K_2C_4H_4O_6 + H_2C_4H_4O_6 = 2HKC_4H_4O_6.$$

Dissolve each powder separately in 50 c.c. of hot water, and pour the normal salt solution into the acid solution in a bottle. The heavy white precipitate is bitartrate of potassium, acid potassium tartrate or more commonly called cream of tartar. When it is cold, filter it off by means of the suction pump and filter flask, wash it once or twice with water, and set it aside to dry. Taste it.

3. Baking powders are mixtures of some kind of a pulverized acid or acid salt with bicarbonate of soda. They generally contain some starch to keep the mixture dry. The best baking powder is one made with cream of tartar. The equation representing its action when moistened is,

$$HNaCO_3 + HKC_4H_4O_6 = NaKC_4H_4O_6 + H_2O + CO_2.$$

From this equation figure out how much bicarbonate of soda should be used to go with 20 grams of cream of tartar. Weigh out the proper quantity. Mix these two powders very thoroughly in a large porcelain mortar or on a large sheet of paper. Add 5 grams of corn starch and mix again. The more thoroughly it is mixed the better the powder will be.

Moisten a small quantity in a test tube and explain the reaction.

4. Write four equations for other kinds of powders using the following acids or acid salts: Tartaric acid, acid potassium sulphate, acid calcium phosphate, $HCaPO_4$, $Al_2(SO_4)_3$. Aluminum sulphate is the commonest ingredient of cheap baking powder. It reacts with the bicarbonate of soda to form sodium sulphate, aluminum oxide, and carbon dioxide. **148.**

LABORATORY EXERCISE NO. 26.

Silicon.

Materials. Water glass; sodium wire; barium chloride sol.; cupric sulphate sol.; sodium carbonate; lead oxide; hydrofluoric acid; sand.

Apparatus. Sealed tubes; evaporating dish; mortar; square of cloth.

1. Heat a small piece of sodium in a sealed tube until there is no further change. Glass is a silicate and contains silicon dioxide. Sodium is a powerful reducing agent. When heated with the glass it takes oxygen from the silicon dioxide and liberates silicon.

2. Break the tube by plunging it while hot into water in the evaporating dish. Decant off the excess of water. The black material is silicon adhering to the pieces of glass. To test this add sodium hydroxide sol. and observe very closely to see if a colorless gas is given off. The gas is hydrogen.

$$4NaOH + Si = Na_4SiO_4 + 2H_2.$$

3. Place about 10 c.c. of water glass in mortar. **Do not measure it out.** Add about one-half as much dil. hydrochloric acid. Wash the jelly with water and strain through a piece of cloth. Transfer the material to the iron plate or the iron crucible and heat as hot as possible.

$$Na_4SiO_4 + 4HCl = 4NaCl + H_4SiO_4 \text{ (jelly)}.$$
$$H_4SiO_4 \text{ heated} = SiO_2 + 2H_2O.$$

The white residue is silicon dioxide or silica. Is it an acidic or a basic oxide?

4. Test the solubility of this material in hydrochloric acid and in sodium hydroxide sol. **64-g.**

5. Dilute the water glass with five times as much water and divide into two parts. To one add barium chloride solution and to the other cupric sulphate sol.

6. The following equations represent the reaction involved in the formation of lead sodium glass.

$Na_2O + PbO + 2SiO_2 = Na_2SiO_3 + PbSiO_3$ or, since the sodium oxide is usually added in the form of sodium carbonate,

$$Na_2CO_3 + PbO + 2SiO_2 = Na_2SiO_3 + PbSiO_3 + CO_2.$$

Weigh out and mix these things, using one gram of the silica you have made, and the proper proportions of lead oxide and sodium carbonate. Grind together in mortar and heat in sealed tube. The product is a variety of glass.

7. Warm a glass plate not hotter than can be borne by the hand. Cover one side thoroughly with beeswax and let the plate cool and the wax harden by itself. Make a drawing on the wax with any sharp point, being sure that the marks go through the wax. Cut two pieces of filter paper just the size of the plate, place them on top of the waxed side and have them moistened with hydrofluoric acid. Leave the paper on the plate for five minutes, then remove it carefully, being particular not to get any of the acid on the hands. Wash the plate, warm it gently and remove the wax with a towel.

The hydrofluoric acid has attacked the silica of the glass according to the following equation:

$$SiO_2 + 4HF = 2H_2O + SiF_4.$$

The SiF_4 is a gas.

LABORATORY EXERCISE NO. 27.

Tin and Lead.

Materials. Granulated tin; stannous chloride solution; mercuric chloride sol.; ammonium sulphide; strips of zinc; lead wire; lead nitrate sol.; potassium chromate sol.

Apparatus. Test tubes; stopper and delivery tube; lamp stand.

1. Dissolve a small piece of tin in warm dilute hydrochloric acid. The gas given off is hydrogen. Fix the tube in clamp of lamp stand with bottom of the tube near the base of the flame as in fig. 12, and let the action continue for at least ten minutes. The solution contains stannous chloride, $SnCl_2$. **120.**

2. To 1 c.c. of mercuric chloride sol. add stannous chloride from reagent bottle drop by drop, until no further change takes place. The white precipitate that forms at first is mercurous chloride, HgCl. The final grey precipitate is metallic mercury. The stannous chloride is changed to stannic chloride, $SnCl_4$, which remains in solution. Write two equations.

3. Repeat this test with the stannous chloride which you have made.

4. To 2 c.c. of stannous chloride solution add 5 c.c. of water, and then hydrogen sulphide. Filter off the precipitate. Stand the funnel in a clean test tube and add a little yellow ammonium sulphide to the filter. Yellow ammonium sulphide contains an excess of sulphur dissolved in it. The brown sulphide dissolves because it forms with the yellow ammonium sulphide a soluble compound known as ammonium thiostannate $(NH_4)_2SnS_3$. This runs through the filter into the test tube. Add some hydrochloric acid to this solution. The following reaction takes place:

$$(NH_4)_2SnS_3 + 2HCl = 2NH_4Cl + H_2S + \underline{SnS_2}.$$

5. To 2 c.c. of stannous chloride solution add ten drops of conc. nitric acid, and heat gently. This oxidizes the stannous chloride solution to stannic chloride. Represent by equation.

Remember that this solution contains an excess of free hydrochloric acid. **77.** Dilute with 10 c.c. of water and pass in hydrogen sulphide. The precipitate is stannic sulphide. Filter it off and treat it just as you did the stannous sulphide.

6. Place about 5 c.c. of stannous chloride solution from reagent bottle in t.t. and add a strip of zinc. **121.**

7. Examine a piece of lead wire. Try to dissolve it in sulphuric, hydrochloric, and hot dilute nitric acid.

8. Test a solution of lead nitrate with sulphuric acid and with potassium chromate solution. Pass hydrogen sulphide into a very dilute solution of lead nitrate.

9. To 2 c.c. of lead nitrate solution add one drop of sodium hydroxide solution. Now add an excess of sodium hydroxide. Compare the corresponding reaction with tin.

10. To 5 c.c. of lead nitrate solution in test tube add a strip of zinc. Leave it for at least ten minutes. **121.**

LABORATORY EXERCISE NO. 28

Sodium and Potassium.

Materials. Sodium bicarbonate; lime water; sodium carbonate; sodium chloride; calcium chloride; barium chloride; sodium nitrate; sodium sulphate; potassium chloride; hydrogen potassium tartrate; potassium nitrate; potassium sulphate; carbon dioxide.

Apparatus. Test tubes; stopper and delivery tube; iron wire with loop on end; glass plate; beaker; wire gauze; lamp stand; copper wire; square of blue glass.

1. Shake up a teaspoonful of common salt in a bottle with a t.t. full of ammonia. Filter off ½ t.t. of the solution. Pass carbon dioxide through the solution for about an hour. Set the tube away until the next laboratory exercise. The precipitate is sodium bicarbonate, made by the Ammonia, or Solvay Process. Filter it off and prove that it is bicarbonate.

2. Heat 5 c.c. of bicarbonate of soda in a test tube having a delivery tube that passes into lime water. **Do not melt the tube.** What is left in the tube? When the tube is cold add 2 or 3 c.c. of cold water. Explain the evolution of heat.

3. Dissolve a very small quantity of common salt in water and let some of the solution evaporate on a glass plate. Examine the crystals under the microscope.

4. Add some calcium chloride solution, also some barium chloride solution, to separate portions of sodium carbonate solution.

5. Clean an iron wire by dipping it repeatedly in a test tube containing conc. hydrochloric acid, washing and heating, until it gives no color when heated in the Bunsen flame. Moisten it with water, touch it to some sodium salt and heat in the Bunsen flame. What color does sodium impart to the flame? Observe this color through the blue glass.

6. Dissolve 8 g. of potassium chloride and 10 g. of sodium nitrate in 20 c.c. of water in beaker on wire gauze on lamp stand, and boil off half of the water. Let the precipitate settle and pour the liquid off into a test tube. Dissolve a very small quantity of the residue in water and let a few drops of the

solution evaporate on a square of glass. Examine the crystals under the microscope. What are they? Allow the tube containing the liquid to stand until cold. Crystals will form. What must they be?

7. Test compounds of potassium in the flame just as you did the sodium. Note the color through the blue glass. Put both compounds of sodium and potassium on the wire at once and note the flame color without and with the blue glass. Which color predominates without the glass? Which with it?

LABORATORY EXERCISE NO. 29.

Calcium, Strontium, and Barium.

Materials. Quicklime in lumps; sodium carbonate; strontium nitrate; barium chloride sol.; calcium chloride.

Apparatus. Lamp stand; beaker; wire gauze; test tubes; iron wire; flask.

1. Heat ½ t.t. of water to boiling in beaker on wire gauze. Add a lump of lime half the size of an egg. The lime ought to crumble to a white, dry amorphous powder. **129.** This is dry slaked lime or calcium hydroxide, $Ca(OH)_2$.

2. Add water and stir to the consistency of milk. This is "milk of lime" or whitewash. Dilute a portion of this very much with water and filter. Feel the filtrate with the fingers, test it with litmus paper and taste it. It is a very dilute solution of calcium hydroxide and is known as lime water. Do you consider calcium hydroxide very soluble?

3. Dissolve about 4 g. of sodium carbonate in boiling water in flask and add a tablespoonful of milk of lime. Stir thoroughly and filter. Feel the filtrate with the fingers. Test it with litmus paper. What is it? What remained on the filter?

4. Dissolve some milk of lime in hydrochloric acid and test the product on iron wire in the flame. Add sodium carbonate solution to some calcium nitrate solution.

5. Make some strontium nitrate solution and add some sodium carbonate solution. To another portion add some sul-

phuric acid. Test some of the dry salt moistened with hydrochloric acid in the flame.

6. Repeat these tests with barium chloride. How could you distinguish between barium and strontium?

LABORATORY EXERCISE NO. 30.

MAGNESIUM, ZINC, CADMIUM AND MERCURY.

Materials. Magnesium wire; hydrogen disodium phosphate; granulated zinc; ammonium sulphide; cadmium sulphate; strips of zinc; mercury; mercuric chloride; hydrogen sulphide; brass.

Apparatus. Bottle; evaporating dish; test tubes.

1. Burn some magnesium wire in a bottle. Don't drop the burning wire into the bottle, but hold it in the forceps while burning. Shake the product with water and test with litmus paper. Is magnesium oxide basic or acidic? What is formed when it is dissolved in water?

2. Add magnesium wire to 2 or 3 c.c. of dilute hydrochloric acid in porcelain dish until no more will dissolve. Dilute the solution. What does it contain? To a portion of this solution add ammonium hydroxide. What is the precipitate? **147.**

3. To another portion add 10 c.c. of hydrochloric acid and an excess of ammonia. What compound is present here which prevents the precipitation of magnesium hydroxide? **147.**

4. To this same solution add a few drops of hydrogen disodium phosphate. The precipitate is ammonium magnesium phosphate, NH_4MgPO_4. This is the common test for magnesium.

5. Make a careful test to see if hydrogen is given off when magnesium is boiled with water.

6. Dissolve about a gram of zinc in 5 c.c. of dilute nitric acid. What have you in solution? Represent the reaction by an equation. What gas was given off? Dilute the solution.

7. To a portion of this solution add sodium hydroxide drop by drop until a precipitate is formed. What is the precipitate? Then add an excess of the hydroxide. Why does the precipitate dissolve? Toward a strong acid, zinc hydroxide acts like a base; toward a strong base it acts like an acid, having the

formula H_2ZnO_2. What soluble compound is formed when zinc hydroxide dissolves in sodium hydroxide? What other hydroxides act like zinc hydroxide in this respect?

8. Pass hydrogen sulphide into 10 c.c. of ammonia in a test tube for five minutes. **142.** What compound is formed? Add this solution to a small quantity of the zinc solution. What is the precipitate? Is this precipitate soluble in dilute hydrochloric acid? **144.** Is copper sulphide soluble in dilute hydrochloric acid? How could you separate zinc and copper? **145.** Dissolve a very small piece of brass and prove that it contains both zinc and copper.

9. Find out what cadmium sulphide is like and whether it is soluble in dilute acids.

10. Dissolve a very small globule of mercury in 2 or 3 c.c. of hot dilute nitric acid. Dilute the solution. Pass hydrogen sulphide into a portion of the solution and dip a bright piece of copper wire into another portion. **121.** The solution contains mercuric nitrate.

11. Cover a small globule of mercury with cold dilute nitric acid and let it stand until the next laboratory exercise. This solution will contain mercurous nitrate, $HgNO_3$. To a portion of it add a few drops of hydrochloric acid. What other chlorides are insoluble? Filter off the precipitate and add ammonia to the filter.

LABORATORY EXERCISE NO. 31.

Copper and Silver.

Materials. Copper wire; silver wire; iron wire; blue vitriol; grape sugar solution; cupric sulphate sol.; potassium iodide sol.; silver nitrate sol.; sodium chloride.

Apparatus. Test tubes.

1. Place about one inch of dilute nitric acid in a t.t. and add as much copper wire as the acid will dissolve. Do this in the hood. Describe and explain the reaction.

2. Divide the solution into two parts. Heat one of these and add ammonia little by little, with much shaking, until the precipitate at first formed dissolves completely. Stand the tube in the rack until the next laboratory exercise. Beautiful blue

crystal oi cupro-ammonium nitrate will be found. Formula, $Cu(NO_3)_2.4NH_3$.

3. Dilute the other portion with water. Pass hydrogen sulphide into a portion of this solution diluted with water. **145.** Pass hydrogen sulphide into some very dilute cupric sulphate solution. See if it is possible to remove all the copper from solution in this way. Heat some blue vitriol in a sealed tube and note the change in the color. Allow some of the water to run back on the anhydrous material.

To 5 c.c. of grape sugar solution in test tube add 5 c.c. of cupric sulphate solution. Now add sodium hydroxide solution, shaking until the precipitate first formed is redissolved.

Warm carefully, noting the changes. Let it stand. The precipitate is cuprous oxide. The grape sugar acted as a reducing agent: do not attempt to make use of its formula in your equation.

4. Put a piece of iron wire into some cupric sulphate solution. **121.**

5. Add some sodium hydroxide solution to some cupric sulphate solution. Heat to boiling. The precipitate first formed is cupric hydroxide. When it is heated each molecule gives up one molecule of water. What is the black precipitate you finally obtain? Solutions of salts of many other heavy metals behave in the same way. See Lab. Ex. 35-4.

6. To a small portion of some copper solution add ammonium hydroxide little by little at first, and finally in excess. This intense color is quite characteristic of copper.

7. Dissolve one inch of silver wire in about one inch of dilute nitric acid. Describe the reaction and write the equation. Add water to half fill the tube and explain later why this makes the solution cloudy.

8. To 2 c.c. of this solution add a solution of sodium chloride. Boil the contents of the tube. **139.** Get the precipitate on a filter and expose it to sunlight. To another portion add a piece of copper wire. **121.**

9. To 1 c.c. of silver nitrate solution add a little potassium iodide solution. Pass hydrogen sulphide through a very dilute solution of silver nitrate.

10. Devise a method for testing for silver, lead and mercurous mercury contained in the same solution.

LABORATORY EXERCISE NO. 32.

Iron and Manganese.

Materials. Green vitriol; fine iron wire; chlorine water; potassium nitrate; red lead; ammonium sulphate; carbon dioxide; cupric sulphate sol.; potassium ferrocyanide sol.; ammonium thiocyanate; potassium ferricyanide; manganese dioxide; sodium nitrite; potassium hydroxide; potassium permanganate; sulphurous acid.

Apparatus. Flask; lamp stand; wire gauze; test tubes; beaker; evaporating dish; iron crucible.

1. Place 40 c.c. of water in the flask, add 10 c.c. of conc. sulphuric acid and mix the liquids by shaking. Weigh out about 10 grams of fine iron wire and place it in the dilute acid. Heat the flask on the lamp until the reaction will proceed by itself, then stand the flask in the hood until all the iron is dissolved. Filter the hot solution and divide the filtrate into two portions. Set one portion aside to crystallize and dilute the other for use in the following experiments.

2. Test the action of the dilute and conc. cold and hot acids on bits of iron wire. See Ex. 18-5.

3. To a few drops of ferrous sulphate sol. from 1, add a little ammonium hydroxide sol. The precipitate is ferrous hydroxide. Watch it for some time and explain what happens to it. Add conc. nitric acid until the precipitate dissolves, then boil. Cool, and again make alkaline with ammonia. What did the nitric acid do to the ferrous sulphate solution? **77.**

4. Make a little ferrous chloride solution by dissolving iron wire in hydrochloric acid, and add chlorine water in excess. An excess is present if an odor of chlorine remains after warming. Add ammonium hydroxide and explain result. What did the chlorine water do to the ferrous chloride? **77.**

5. To a few drops of ferric chloride sol. add ammonium hydroxide. Note the color of the precipitate. It is ferric hydroxide. Dissolve it in hydrochloric acid and add a bundle of fine iron wire. Heat to boiling from time to time for three minutes. Pour some of the almost colorless solution into hot ammonia. What is the precipitate? And what did the iron wire do to the ferric chloride? **81.**

6. Add a few drops of ammonium sulphide to a little ferrous sulphate solution. To a little ferric chloride solution add ammonium sulphide and acidify with hydrochloric acid. What is the undissolved residue? Make the solution alkaline with ammonium hydroxide and explain what the ammonium sulphide must have done to the ferric chloride. **82.**

7. Dissolve 20 g. of green vitriol, crystallized ferrous sulphate, $FeSO_4 + 7H_2O$, and the proper quantity of ammonium sulphate to make ferrous ammonium sulphate, $(NH_4)_2SO_4.FeSO_4 + 6H_2O$, in 50 c.c. of water, boiling, in flask, and filter into bottle. Set aside to crystallize. How would you test these crystals for iron, ammonium and for sulphate?

8. Test both ferrous and ferric iron solutions with each of the following reagents and tabulate the results: Potassium ferrocyanide, potassium ferricyanide, and ammonium thiocyanate. Devise a method for testing for ferrous and ferric iron, each in the presence of the other.

9. Oxidize a little potassium ferrocyanide solution with aqua regia. **77.** Dilute the solution and add ferrous sulphate. Explain the result.

10. What is the action of ammonium hydroxide, also of ammonium sulphide, on potassium ferrocyanide sol.? Explain.

11. Heat a small quantity of manganese dioxide in a sealed tube. What gas is given off, and what is left in the tube? Heat ¼ g. of manganese dioxide in a t.t. with 5 c.c. of conc. HCl. What gas is given off? What kind of a reaction is this?

12. Shake up some manganese dioxide with a strong solution of sulphur dioxide in water. Add a small piece of sodium nitrite to a tube containing one-fourth gram of manganese dioxide and 5 c.c. of dilute nitric acid. Heat to boiling. Add one-half gram of red lead. Some of the manganese is oxidized to permanganic acid.

13. Fuse one gram of manganese dioxide with one gram of potassium hydroxide and a small quantity of KNO_3, in iron crucible. Dissolve out the product with water, filter, and pass carbon dioxide into the solution. Potassium permanganate is formed in solution. **77.**

14. Dilute some of the ferrous sulphate solution with water, and add potassium permanganate solution. **77.**

LABORATORY EXERCISE NO. 33.

Aluminum and Chromium.

Materials. Aluminum wire and foil; alum; borax; sodium carbonate; aluminum sulphate; ammonium sulphate; ammonium sulphide; chrome alum; potassium chromate; alcohol; lead nitrate; barium chloride; silver nitrate; mercuric chloride sol.

Apparatus. Lamp stand; bottles; test tubes; evaporating dish; beaker; wire gauze.

1. Test the solubility of aluminum in nitric, hydrochloric and sulphuric acids, both dilute and concentrated, also in sodium hydroxide solution. Evaporate the solution of aluminum chloride you have made to dryness in the hood. Heat the residue. When it has cooled try to dissolve it in water.

2. Amalgamate a strip of aluminum foil in mercuric chloride solution, allowing the reaction to continue for several minutes. Wipe off the solution with clean filter paper. **121.**

After a time notice the remarkable change it has undergone and feel it to note the heat of reaction. **68.**

3. To an alum solution add ammonium hydroxide solution until sufficient is present to dissolve the precipitate which first forms. Boil a little of this clear solution. The hydroxide is reprecipitated. This is the regular test for the aluminum radical in solution.

4. Put a little alum solution in each of three test tubes. Add borax solution to one, sodium carbonate solution to another and sodium hydroxide to the third. Compare the results and explain. **148.**

5. Take 10 g. of aluminum sulphate and the required quantity of ammonium sulphate to make ammonium alum. Dissolve the two salts together in 50 c.c. of boiling water. Filter the solution into a bottle and set aside to crystallize. After several days crystals of ammonium alum will be found. Formula, $(NH_4)_2SO_4 . Al_2(SO_4)_3 + 24H_2O$. Get these out, dissolve them in water and test the solution for ammonium, sulphate and aluminum radicals.

6. Add a few drops of ammonium sulphide to an alum solution. Filter off the precipitate and test it for a sulphide. **148.**

7. Add ammonium hydroxide to a little chrome alum solution. Compare the precipitate with that formed by treating common alum in the same way. Treat another portion of chrome alum solution with sodium hydroxide and observe its action with an excess of the precipitant.

8. Precipitate chromic hydroxide with sodium hydroxide and then add enough of the latter to redissolve the precipitate. Dilute the solution with an equal volume of water and then pour away all but about an inch in a test tube of the solution. Add about 5 g. of chloride of lime and boil for several minutes until the solution is quite yellow. Filter into a test tube and acidify the filtrate with hydrochloric acid. Note and explain the successive color changes.

9. Dissolve 10 g. of potassium bichromate in 50 c.c. of water in flask. When solution is complete add 10 c.c. of conc. sulphuric acid. Transfer a few drops to a test tube and add ammonia. To the rest of the solution, under the hood, add 5 c.c. of alcohol. The mixture of sulphuric acid and potassium bichromate oxidizes the alcohol to aldehyde (C_2H_4O), which has a strong irritating odor. Add ammonia to a few drops of this solution in a test tube. Leave the rest of the solution in a bottle to crystallize.

10. Test potassium bichromate solution with lead nitrate, barium chloride and with silver nitrate.

LABORATORY EXERCISE NO. 34.

Review Laboratory Work.

I. The Identification of Common Chemical Compounds.

Preliminary Examination.

1. If it is a solid, pulverize it and put it in a bottle or tube marked with its letter or number.
2. If you recognize it or think you do, make the final confirmatory tests as called for in these directions at once.
3. Test its action toward moist litmus paper.
   a. Turns red, shows an acid anhydride, acid or acid salt. Or salt of a strong acid with weak base.
   b. Turns blue. Shows an alkali, an alkaline oxide or an alkaline carbonate or alkaline salt of weak acid.

4. Heat a small portion in a sealed tube as hot as possible.
   a. There may be no change, possibly a sulphate.
   b. It may give off water, showing water of crystallization or water of decomposition.
   c. It may give off a colorless gas which turns lime water white, showing a carbonate or bicarbonate.
   d. It may give off red fumes, showing a nitrate or nitrite.
   e. It may leave a black residue of carbon, showing it to be an organic compound.
   f. It may simply melt, showing nothing.
   g. It may give a sublimate of sulphur, showing it to be a thiosulphate.
   h. It may volatilize completely, showing it to be either ammonium nitrate or ammonium carbonate.
   i. It may sublime completely, showing it to be an ammonium compound or a mercury compound.
   j. The substance may be yellow when hot and white when cold, showing it to be a zinc compound.
5. Moisten a small quantity of the solid with dil. HCl.
   a. Effervescence of a colorless gas that turns lime-water white shows a carbonate or bicarbonate.
   b. Effervescence of a colorless gas that smells of sulphur dioxide shows a sulphite or a thiosulphate.
   c. Odor of hydrogen sulphide shows presence of a sulphide.
   d. Odor of chlorine shows the substance to be an oxidizing agent.
6. Clean the iron wire loop by dipping it in a t.t. of hot dil. hydrochloric acid, washing and heating until it does not color the flame. Moisten a very small quantity of the powder on a glass plate with conc. HCl. Place a small quantity of this mixture on the clean end of the wire and heat it carefully in the edge of the flame.
   a. Yellow color shows sodium. Remember that sodium is almost always present in small quantities as an impurity.

   Yellow flames should always be examined through the blue glass. A red color when seen through the blue glass shows potassium.

b. A purple color appearing red through the blue glass shows potassium. The blue glass removes the yellow color due to sodium which might obscure the purple.

c. Red color without the glass shows lithium, calcium or strontium.

d. Green color shows barium or copper.

7. Heat about 3 s.s. of the dry powder in a t.t. with 3 drops of water and 3 c.c. of conc. sulphuric acid.

a. No change; possibly a sulphate.

b. Evolution of hydrochloric acid gas; a chloride.

c. Evolution of nitric acid, a colorless liquid distilling up on the sides of the tube, shows a nitrate.

d. Evolution of sulphur dioxide without the formation of sulphur shows a sulphite.

e. Evolution of sulphur dioxide with the formation of sulphur shows a thiosulphate.

f. Bromine distilling up the sides of the tube accompanied by acid fumes shows a bromide.

g. A purple vapor shows an iodide.

h. Evolution of an acid gas that etches the inside of the tube shows a fluoride.

i. Odor of vinegar shows an acetate.

8. Add a small quantity of sodium hydroxide solution to 2 s.s. of the dry material in the mortar. Mix thoroughly with the pestle. The odor of ammonia shows an ammonium compound.

9. If the material is an acid or an acid anhydride it will unite readily with sodium hydroxide solution to form a salt with the evolution of much heat.

10. If the material is a base or a basic oxide it will unite readily with dilute nitric acid to form the corresponding salt with the evolution of much heat.

The above preliminary tests should have given some indication as to the nature of the substance. In case they have, proceed at once to make the final confirmatory tests for the radicals whose presence is suspected. If they have not, the next step is to make a solution of about 5 g. of the material either in water or dilute nitric acid. Make up the solution to 100 c.c. and use small portions of it in the following tests.

Wherever a definite reaction is obtained leave a blank line in the note book to be filled in later with an equation representing that reaction, when the nature of the substance is definitely known.

II. Tests to be made with a Water or Nitric Acid Solution of the Material.

These tests must invariably be made in the order given.

1. Add a few drops of dil. HCl to a small portion of the solution.
   a. A white ppt. shows silver, lead or mercurous mercury.
2. Pass H_2S through a small quantity of the solution to which a few drops of hydrochloric acid have been added if it was not already an acid solution.
   a. A black ppt. shows lead, copper, bismuth or mercuric mercury.
   b. A light yellow ppt. shows cadmium or arsenic.
   c. An orange yellow ppt. shows antimony.
3. Make alkaline with ammonia, and pass H_2S through the liquid.
   a. A black ppt. shows nickel, cobalt or iron.
   b. A pink ppt. shows manganese.
   c. A white ppt. shows zinc.
4. Add ammonia and heat to boiling.
   a. A white gelatinous ppt. shows aluminum, bismuth, magnesium, or a phosphate.
   b. A brown gelatinous ppt. shows ferric iron.
5. Add a solution of sodium carbonate.
   a. A white ppt. shows calcium, strontium, barium or magnesium.
6. To a small portion of the water solution add silver nitrate.
   a. A white ppt. shows chlorine, nitrite or thiosulphate.
   b. A cream colored ppt. shows a bromide.
   c. A light yellow ppt. shows an iodide.
   d. A bright yellow ppt. shows a phosphate.
   e. A deep red ppt. shows a chromate.
7. To a nitric acid solution or solution to which a few drops of nitric acid have been added, add silver nitrate.
   a. White ppt. soluble in ammonia shows chlorine.

b. Cream colored ppt. shows a bromide.
c. Light yellow ppt. shows an iodide.

8. Add barium chloride sol.
   a. White ppt. sol. in dil. HCl shows a phosphate or carbonate.
   b. White ppt. insol. in dil. HCl shows a sulphate.
   c. Light yellow ppt. shows a chromate.

III. Confirmatory Tests.

In each of these tests use small portions of a water or nitric acid solution.

Lead. With HCl, white ppt. sol. in hot water; with potass. dichromate, yellow ppt.; with H_2S, black ppt.; with sheet zinc, black ppt. of metallic lead.

Silver. With HCl, white curdy ppt. sol. in ammonia; with H_2S, black ppt.; with copper, white ppt. of metallic silver.

Mercurous mercury. With HCl, white ppt. that is blackened with ammonia; with H_2S, black ppt.; covers bright copper wire with mercury.

Copper. With excess of ammonia, deep blue sol.; H_2S, black ppt.; deposits copper on bright iron wire; boiled with excess of sodium hydroxide gives black ppt. of cupric oxide.

Mercuric mercury. With stannous chloride gives at first a white ppt. which turns grey on adding an excess; with H_2S, black ppt.; deposits mercury on bright copper wire.

Bismuth. White gelatinous ppt. with ammonium hydroxide. If this ppt. be collected on a filter and a solution of stannous chloride in sodium hydroxide be added to it, it will be blackened.

Cadmium. With hydrogen sulphide gives a yellow ppt. With sodium hydroxide a white gelatinous ppt.

Antimony. With hydrogen sulphide an orange yellow ppt. of antimony trisulphide, completely soluble in yellow ammonium sulphide.

Ferrous Iron. With ammonia a light green ppt. which rapidly turns brown on the surface. With potass. ferricyanide a deep blue ppt.

Ferric Iron. With potassium or ammonium thiocyanate a deep red solution. With ammonium hydroxide a brown gelatinous ppt. With potass. ferrocyanide deep blue ppt.

Aluminum. With ammonia, white gelatinous ppt. soluble in excess of ammonia, insoluble in boiling ammonia.

Basic chromium. When treated with chloride of lime, or sodium oxychloride and sodium hydroxide it is oxidized to chromate and recognized by its yellow color.

Zinc. Hydrogen sulphide passed into ammonia and the solution added to a neutral solution of a zinc salt gives a white precipitate of zinc sulphide.

Ammonium hydroxide gives a white ppt. of zinc hydroxide easily soluble in an excess of ammonia.

Nickel. Sodium hydroxide gives an apple green ppt. insoluble in excess of sodium hydroxide.

Manganese. Fused with a small quantity of sodium carbonate and KNO_3 on platinum foil over blowpipe gives a green color.

Magnesium. Ammonium hydroxide gives a white ppt. of magnesium hydroxide soluble in excess of HCl, from which solution it is not reprecipitated by ammonia. Sodium phosphate added to this solution gives a white ppt. of ammonium magnesium phosphate.

Barium. White ppt. with K_2SO_4 insoluble in HCl. Yellow ppt. with potass. chromate. Green flame coloration.

Strontium. Red flame coloration. White ppt. with potassium sulphate.

Calcium. Yellowish red flame coloration. No ppt. in very dilute sol. with K_2SO_4; white ppt. with Na_2CO_3 sol.

Sodium. Bright yellow flame coloration.

Potassium. Violet flame coloration which appears red through a blue glass. The sodium color obscures the potassium color, therefore all yellow colors must be examined through the blue glass.

Chloride. Silver nitrate gives a white curdy ppt. insoluble in dil. nitric acid and soluble in ammonia.

Bromide. Silver nitrate gives a cream-colored ppt. insoluble in dil. nitric acid. The original sol. treated with a few drops of conc. nitric acid and shaken with a little carbon disulphide gives a yellow or a red globule.

Iodide. A few drops of conc. nitric acid added to the sol. give a ppt. of iodine, which, on boiling, turns into purple vapors.

Sulphide. The dry substance treated with a little dil. sulphuric acid in a t.t. gives off hydrogen sulphide, which will blacken filter paper moistened with lead nitrate and ammonia; or the H_2S gas passed into a $CdSO_4$ sol. gives a yellow ppt.

Thiosulphate. The dry substance treated with sulphuric acid gives off sulphur dioxide, and the solution treated with the acid gives a white ppt. of sulphur as well as an odor of sulphur dioxide. Silver nitrate gives a white ppt. of silver thiosulphate.

Sulphite. Dil. sulphuric acid on the dry substance gives sulphur dioxide; with the solution it gives an odor of sulphur dioxide without a ppt.

Sulphate. Barium chloride gives a white ppt. insoluble in dil. HCl.

Chromate. Lead nitrate gives a yellow ppt. of lead chromate.

Nitrite. Dil. hydrochloric acid gives an evolution of red oxides of nitrogen.

Nitrate. Concentrated sulphuric acid with copper on the dry salt gives nitric oxide.

Phosphate. Silver nitrate in the neutral or water solution gives a yellow ppt. of silver phosphate. Add silver nitrate to the nitric acid solution and then add ammonia carefully. A yellow ring will form in the neutral zone.

Carbonate. The substance effervesces with acids, giving off carbon dioxide, which will give a white ppt. in lime water.

Acetate. The dry material heated with a little conc. sulphuric acid and alcohol gives the odor of ethyl acetate.

Tartrate. Add a few drops of silver nitrate and then ammonia until the ppt. which forms at first is just redissolved. Heat to boiling for some time. A deposit of metallic silver in the form of a mirror will form on the inside of the test tube.

Bicarbonate. Gives water and CO_2 when heated in sealed tube.

LABORATORY EXERCISE NO. 35.

THE PREPARATION OF CHEMICAL COMPOUNDS.

In this exercise the more advanced student is expected to select some particular compound that he desires to make on a larger scale and study out the details of the operation from the following general directions. When he knows exactly what he ought to do to obtain the result, he should be provided with the necessary materials and apparatus.

GENERAL DIRECTIONS FOR CHEMICAL PREPARATIONS.

Acids. Oxygen acids may be made by oxidizing the required non-metal with any suitable oxidizing agent to the corresponding acid anhydride and dissolving this in water.

Or, if they are volatile and stable, by treating their salts with sulphuric acid.

Basic oxides.

1. By heating the metal in oxygen.
2. By heating the carbonate.
3. By heating the nitrate.
4. By heating the hydroxide.

Salts.

1. By treating a metal with an acid.
2. By treating a basic oxide with an acid.
3. By treating a base with an acid.
4. By treating an acid with a salt of a more volatile acid.
5. By double decomposition with the formation of an insoluble compound.
6. By melting together a basic with an acidic oxide.
7. By treating an alkali with an acid anhydride.

Preparation of salts. Before undertaking the preparation of a chemical compound of any description it is necessary to look up its properties, especially its solubility, and the best

method of making it. If it is desired to make a chemically pure compound it is desirable to use chemically pure materials.

Insoluble salts. Suppose it is desired to make a compound that is virtually insoluble in water. It is commonly made by mixing two solutions, one of which contains the positive radical and the other the negative radical of the required substance.

Select two chemically pure and soluble compounds and calculate exactly how many grams of each will be necessary to form the desired amount of the insoluble compound, being sure to take into account any water of crystallization that they may contain.

Look up their solubilities in water and dissolve each in hot water and add enough cold water to make a cold saturated solution.

Filter the solutions and add one solution to the other very slowly with violent stirring.

Test a small portion on a suction filter to see if it can be filtered and washed. If it is found possible to filter it in this way, wash it rapidly by decantation, pouring the washings through a large suction filter, finally transfer all the material to the filter and wash by pouring water through it until the washings show no test for the negative radical of the other compound formed.

If it is impossible to filter and wash it on the filter let it settle in a narrow deep vessel over night and siphon off the wash water the next day. Add more water, stir thoroughly, let settle and siphon again. Repeat this as many times as necessary to get rid of the other compound present; or it might be filtered through an ordinary folded filter.

Soluble salts. 1. Look up the solubility of the salt at 100° C. and at 20° C. The difference between these will be approximately the amount of the salt that will crystallize out of 100 c.c. of a hot saturated solution when it cools down to 20° C. If it is desired to have 100 g. crystallize out, as many hundred c.c. of a hot saturated solution will be necessary as this difference is contained in 100.

Having found the volume of the saturated solution, find how many grams of the substance will have to be made to saturate it. From this compute the quantities of the materials necessary.

The final solution is filtered hot and the hot filtrate, in a beaker or flask well covered, is thoroughly wrapped up in cloth to cause it to cool very slowly, thereby forming large crystals, and left to stand over night.

If pure materials have been used the product is essentially chemically pure. In any case it may be recrystallized several times if necessary to render it more pure.

To find the number of cubic centimeters of acid to use, the theoretical weight of the acid is divided by the quantity of acid in 1 c.c. of the solution. If much dilute acid is to be used it will be necessary to take the water it contains into account in computing the quantity of water to add to make up the required volume.

Example. Suppose it is desired to make potassium nitrate and have 100 g. crystallize out from solution. 100 c.c. of water at 100° C. dissolve 247 g. of the salt, and at 20° C., 31 g.; therefore on cooling such a saturated solution to 20° C. 216 g. ought to crystallize out. Leaving out of account the change in the volume due to the presence of the salt, the volume of water which would cause 100 g. to separate would be $\frac{100}{216} \times 100$ c.c. which equals 46 c.c.

1 c.c. of a saturated solution at 100° C. contains 2.47 g. of the salt. Therefore the total quantity to be made in this case is 2.47×46, which equals 114 g.

Suppose the salt is to be made by mixing potassium carbonate with dilute nitric acid. Then according to the following equation:

$$\underset{138}{\overset{78\text{ g.}}{K_2CO_3}} + \underset{126}{\overset{71\text{ g.}}{2HNO_3}} = \underset{202}{\overset{114\text{ g.}}{2KNO_3}} + H_2O + CO_2.$$

$\frac{138}{202} \times 114 = 78$ g. of potassium carbonate and $\frac{126}{202} \times 114 = 71$ g. of nitric acid will be necessary.

The specific gravity of the acid is 1.28 and it contains 45.5% nitric acid. One c.c. of it therefore contains .58 g. of HNO_3 and .70 g. water. The number of c.c's. necessary would be, $\frac{71}{58} = 24$ c.c. Since one c.c. of this acid contains .70 c.c. of

water the whole will contain 86 c.c., which is more than that required to dissolve the salt formed. No more water need therefore be added. The dilute acid is added little by little to the carbonate in a flask. The heat of reaction will probably be sufficient not only to heat the solution to boiling but to evaporate some of the excess of water. **129.**

The preparation of the hydroxides.

1. In the case of the alkalies, add milk of lime to a solution of the carbonate.
2. In the case of the less soluble hydroxides, add a very soluble hydroxide to a solution of one of the salts.

The preparation of acid salts. These can only be prepared from polybasic acids. Heat the normal salt with the reauired excess of acid.

QUESTIONS ON THE LABORATORY EXERCISES.

Exercise No. 1.

What are the essential elements of a Bunsen burner? Make a drawing showing the essential construction. In what four ways may a burner be found burning? Which of these is the proper way? What are the two ways that are decidedly wrong? How may the wrong ways be easily corrected? Which way is undesirable although harmless? Describe the flame. Which is the hottest part?

Exercise No. 2.

Describe how to cut glass tubing. In rounding the edges of a glass tube why should the other end be held lower down? Why does the flame appear yellow when the glass is heated? In making a sealed tube why is it undesirable to leave a great mass of glass on the end? Why cannot glass tubing be properly bent in the Bunsen flame?

Exercise No. 3.

Define the terms, acid; alkali; basic oxide; alkaline oxide and acidic oxide. What is the effect of thrusting a spark into a bottle of oxygen? If the wood contains carbon and hydrogen,

what products would be formed? **68.** What is the product formed when magnesium burns in oxygen? How does this moist product affect litmus paper? In what will basic oxides generally dissolve? State some characteristic properties of acids.

Exercise No. 4.

Define the term "Properties of matter." What topics should be considered when describing the properties of a given kind of matter? See outline for recitation. Describe some properties of sodium. Describe the visible effects of putting sodium on wet filter paper. Is sodium oxide a basic or acidic oxide? What would be formed if it were dissolved in water? What is the effect of heating mercuric oxide? What gas comes out of the tube? What is left in the tube? What is left in the tube after heating maganese dioxide? **71.** What is the effect of heating red oxide of lead? What kinds of oxides dissolve readily in acids?

Exercise No. 5.

Why is a mixture of manganese dioxide and potassium chlorate used for preparing oxygen? What is left in the tube after no more oxygen can be obtained? How could you prove this? Why does the oxygen look smoky? What compound is supposed to be formed when carbon dioxide dissolves in water? What kind of an oxide is CO_2, acidic or basic? What two oxides does phosphorus form? Under what conditions would each result? What acids would they form with water? What oxides does sulphur form? What acids would these oxides form with water? What gases would burn in oxygen? Would oxygen burn in any gas?

Exercise No. 6.

Why should the bottles be kept inverted? What would be the color of the flame if the gas issued from a platinum tube? Why is it yellow when burning on the end of a glass tube? If the gas dissolved in water, why should there be a suction on the hand? What is the most explosive mixture of hydrogen and air? Why should the hydrogen coming directly from a generator not be ignited? Why was water put in the apparatus? The

action might stop for three different reasons, what are they? State three ways in which hydrogen might be obtained from water. Two ways in which it might be obtained from acids. When sulphuric acid acts on zinc, what else is formed? Describe three different experiments proving that hydrogen is lighter than air. Describe one experiment proving that water is formed when hydrogen burns.

Exercise No. 7.

What is meant by electrolysis? What is an electrolyte? Which is the positive electrode? Which the negative? Will distilled water allow the current to pass through it? Will hydrant water allow the current to pass? What effect has the presence of a small quantity of sulphuric acid on the conductivity of distilled water? What are the visible effects of passing a current through a solution of potassium sulphate colored with litmus solution?

Exercise No. 8.

What is meant by reduction? What is a reducing agent? What is a precipitate? When is a substance said to be crystalline? When amorphous? What is water of crystallization? What is meant by efflorescence? What is meant by distillation and what is the distillate? What is the effect of heating a metallic oxide in an atmosphere of hydrogen? What class of impurities does filtration remove from water? What kinds of impurities does distillation remove?

Exercise No. 9.

What is meant by a cold saturated solution; a hot saturated solution; a super-saturated solution? How are such solutions made? How may a super-saturated solution be caused to crystallize? What is water of crystallization? What is the complete formula of Glauber's Salt?

Exercise No. 10.

What is an acid; a base; a salt; a basic oxide? What is the action of an acid on a metal? What gas is generally formed? Under what circumstances is this gas not set free?

How many grams of hydrochloric acid would be necessary to exactly combine with ten grams of iron? **151.** What is green vitriol? What is formed when magnesium dissolves in hydrochloric acid? What is the action of an acid on a basic oxide? What is always formed in this kind of action? How many grams of zinc oxide would be necessary to form ten grams of zinc sulphate? Write equations showing the formation of any salt in two different ways.

Exercise No. 11.

What is an alkali; a base? What is the action of an acid on a base? What is first formed? What else is formed? When is a liquid said to be neutral? What is meant by decant? In 1, just before the solution became neutral there was a slight effervescence, to what was it due? Why was it desirable to have the solution slightly acid rather than slightly alkaline? When is one liquid said to be more volatile than another? What is the effect of mixing an acid with a salt of a more volatile acid? If the more volatile acid is unstable, what products may be given off? Name some non-volatile acids; some acids that are volatile and stable; some acids that are volatile and unstable. What is the effect of adding an acid to a carbonate? How many grams of sodium carbonate would be necessary to form ten grams of sodium chloride?

Exercise No. 12.

State the full solubility rule. **166.** What is the effect of mixing two solutions one of which contains the basic radical and the other the acidic radical of an insoluble compound? How did you make barium sulphate? Describe lead chromate, calcium carbonate and barium carbonate. What is formed when a basic oxide is fused together with an acidic oxide? In 1, why should the filtrate be acidified with hydrochloric acid? Why should it effervesce on the addition of the acid? If five grams of sodium carbonate had been used, how many grams of calcium chloride should have been necessary to go with it in order that there should not be an excess of either material? Name the insoluble sulphates; the insoluble chlorides. State the general conditions under which precipitates are formed.

129. Any oxygen salt may be considered as made up of what two kinds of compounds? Write equations showing the formation of any oxygen salt in six different ways.

Exercise No. 13.

Represent all reactions involved in this exercise by equations. How may the hydrogen be removed from hydrochloric acid? Name ten oxidizing agents and state how they oxidize. How many grams of manganese dioxide would be necessary to form ten grams of chlorine? In 2, why is the color of the escaping gas reddish brown? What acid should theoretically be formed when hydrochloric acid is added to potassium chlorate? If this acid were volatile and unstable, what products might be given off? In a mixture of several salts, if sulphuric acid were present in excess, what salts would remain after heating? **64-d.** What kind of colors does chlorine bleach? When nitric acid is added to ammonium hydroxide, what salt is formed? In 8, what is the final precipitate?

Exercise No. 14.

Why must the funnel tube go to the bottom of the flask? Why is water put in the flask? Why is hydrogen chloride visible in moist air? What are the white fumes in D? Represent all reactions involved in this exercise by equations. How many grams of hydrogen chloride could be made from twenty grams of salt? How many liters would this gas occupy at normal temperature and pressure? **152. 105.** State all the evidence you have that the liquid in D is the same as the hydrochloric acid in the reagent bottle. How soluble is hydrogen chloride gas in water? In 5, what makes the water rise into the upper bottle? How might acids other than hydrochloric be made? Write equations showing the formation of hydrochloric acid from ten different chlorides. What would be formed if chlorine were allowed to act on wet phosphorus? **71, 73**

Exercise No. 15.

How might bromine be obtained from hydrobromic acid? What oxidizing agents would remove the hydrogen from hydrobromic acid? Is hydrobromic acid more or less stable

than hydrochloric acid? What takes place when chlorine water is added to a solution of a bromide? Is hydriodic acid more or less stable than hydrobromic acid? Could this acid be made by the action of sulphuric acid on an iodide? Is iodine volatile with steam? How may iodine be separated from an iodide? What is the effect of adding either chlorine or bromine water to a solution of an iodide? Describe tests for both free and combined iodine. What is the effect of passing hydrogen sulphide into iodine suspended in water? What would be formed if phosphorus were added to iodine water? Name three good solvents for iodine. Represent all the reactions involved in this exercise by equations. How many grams of phosphorus would be necessary to form ten grams of hydriodic acid?

Exercise No. 16.

Represent all the reactions involved in this exercise by equations. Draw a diagram of the crystals obtained by the evaporation of the carbon disulphide solution. Describe all the changes that sulphur undergoes when very slowly heated. Describe three different varieties of sulphur. How would you change any one of them into any other? What is formed when sulphur burns in air? What is formed when this oxide is dissolved in water? What is the effect of adding an oxidizing agent to this solution? How many grams of ferrous sulphide could be made from ten grams of iron? How many liters of hydrogen sulphide could be obtained from this amount of ferrous sulphide?

Exercise No. 17.

Represent by equations all the reactions involved in this exercise. How would you obtain hydrogen sulphide from sulphur? How many grams of iron would be necessary to produce ten liters of hydrogen sulphide at normal temperature and pressure? What is formed when hydrogen sulphide burns? What evidence have you that hydrogen sulphide water has acid properties? State three ways in which sulphur dioxide may be made. How could you make sodium sulphite? Show the formation of sulphur dioxide by the oxidation of metals by

means of sulphuric acid. How many grams of copper would be necessary to form ten liters of sulphur dioxide? Describe the preparation of sulphuric acid. How can you prove that nitric acid changes sulphurous acid to sulphuric acid?

Exercise No. 18.

Represent all the reactions involved in this exercise by equations. How many grams of sulphur would have to be burned to form ten grams of sulphuric acid? How many grams of bromine would theoretically be necessary to oxidize ten grams of sulphur as sulphurous acid to sulphuric acid? Describe the test for the sulphate radical. How would you distinguish between a soluble phosphate and a sulphate? Give the formulas of ten oxidizing agents and state how they oxidize.

Exercise No. 19.

Represent all the reactions of this exercise by equations. What two oxides does phosphorus form? Under what circumstances is each formed? How does burning phosphorus over water prepare nitrogen? Does it really prepare it? What becomes of the white smoke? What acids does phosphorus form? What proportion of the air is nitrogen? Describe a method for preparing chemical nitrogen as distinguished from atmospheric nitrogen. What is formed when an ammonium compound is heated? What might be put with an ammonium compound in order to hold the acid in combination and allow the ammonia to escape? Describe the tests for both free and combined ammonia. How many grams of ammonium chloride would be necessary to form ten liters of ammonia?

Exercise No. 20.

Represent by equations, all the reactions involved in this exercise. How many liters of nitrous oxide could be prepared from ten grams of ammonium nitrate? How could you distinguish nitrous oxide from oxygen? How many grams of copper would be necessary to prepare ten liters of nitric oxide? What acid is theoretically formed by the addition of an acid to a nitrite? Describe the preparation of nitric acid. What is a good test for nitric acid?

Exercise No. 21.

1. A substance melts when heated in the sealed tube, colors the flame yellow, gives a white curdy precipitate with silver nitrate which is soluble in ammonium hydroxide, gives a colorless gas that fumes in air when treated with conc. sulphuric acid; what is it?

2. A substance when heated in a sealed tube gives off red fumes, the residue melts and solidifies to a yellow solid when cold. It gives a reddish color to the flame; it is soluble in water. The water solution gives a black precipitate with hydrogen sulphide which does not dissolve in hydrochloric acid, a yellow precipitate with potassium chromate, a white precipitate with sulphuric acid. The solid treated with conc. sulphuric acid gives off a colorless liquid that distills up on the side of the test tube. A piece of copper placed in this mixture causes an evolution of red fumes. What is the substance?

Exercise No. 22.

Represent by equations, all the reactions involved in this exercise. Give the names of all the elements of the fifth group. What oxides do they form? Which are basic and which acidic? What acid does phosphorus trioxide form with water? Name three acids that phosphorus pentoxide forms with water. In 2, why does the phosphorus take fire so readily? What is formed when red phosphorus is dissolved in nitric acid? **73.** How many liters of nitric oxide would be formed if ten grams of phosphorus were dissolved in nitric acid? What oxides does arsenic form? Are these oxides basic or acidic? What acids would these oxides form with water? Describe two different forms of arsenic. What is formed when arsenic burns in air? What is formed when arsenic is dissolved in nitric acid? What is formed when arsenic trioxide dissolves in alkalies? What acids would antimony form?

Exercise No. 23.

Represent by equations, all the reactions involved in this exercise. What are some of the products formed when wood and paper are heated in a sealed tube? What classes of impuri-

ties does charcoal remove from solution? How many liters of carbon dioxide would be formed if ten grams of lead oxide were reduced to lead by charcoal? What are some of the products formed when soft coal is heated in a sealed tube? How is coke made? How is charcoal made? Write equations showing the reduction of metallic oxides with carbon.

Exercise No. 24.

Represent by equations, all the reactions involved in this exercise. What oxides does carbon form? What oxide does carbon form when burned in an excess of oxygen? How many grams of carbon would be necessary to form ten liters of carbon dioxide at normal temperature and pressure? State three ways in which carbon dioxide may be prepared. What acid does this oxide form with water? Is it a stable or unstable acid? What are the black specks which are formed when magnesium burns in a mixture of carbon dioxide and air? What is formed when carbon dioxide is passed into lime water? What are the two products which would be formed if carbon dioxide were passed in excess into sodium hydroxide solution? What is formed when sodium bicarbonate is heated? Describe some properties of calcium chloride

Exercise No. 25.

What is a polybasic acid; a dibasic acid; a tribasic acid; an acid salt; a normal salt, and a basic salt? In 1, what proof have you that the product was hydrogen sodium carbonate? What is the essential composition of a baking powder? Write equations showing the action of several different kinds of baking powders. What salt is the product of the action of a cream of tartar powder?

Exercise No. 26.

Represent by equations, all the reactions involved in this exercise. What product is formed when sodium is heated with silicon dioxide? What is the action of silicon on sodium hydroxide? What acids does silicon form? How could you make silicon dioxide from sodium silicate? Is silicon dioxide a basic or an acidic oxide? State three ways of dissolving

sand. What does glass consist of? What is the action of hydrofluoric acid on sand? How many liters of silicon tetrafluoride would be formed by the action of hydrofluoric acid on ten grams of silicon dioxide?

Exercise No. 27.

Represent by equations, all the reactions involved in this exercise. What are the elements of the fourth group? What oxides do tin and lead form? Are these oxides basic or acidic? What acids does tin form? What salt is formed when tin is dissolved in hydrochloric acid? How many liters of hydrogen would be formed if ten grams of tin were dissolved? Explain the changes that take place when stannous chloride is gradually added to a solution of mercuric chloride. What is the color of stannous sulphide? When this is dissolved in yellow ammonium sulphide, what compound is formed? When this solution is acidified with hydrochloric acid, what is the precipitate? What is the color of stannic sulphide? How could you change stannous sulphide to stannic sulphide? Explain the effect of adding a piece of zinc to a solution of stannous chloride. **121.** What will dissolve lead? In what acids will it not dissolve? Describe lead sulphide, lead sulphate, lead chromate. How many grams of zinc would precipitate ten grams of lead from solution? How many grams of lead would produce ten liters of nitric oxide?

Exercise No. 28.

Represent by equations, all the reactions involved in this exercise. What reactions are involved in the Solvay Process? How would you form calcium and barium carbonates? How do sodium and potassium color the Bunsen flame? Which color would obscure the other? How is it possible to see the potassium color in the presence of sodium? How is potassium nitrate made? How many grams of bicarbonate of soda would be necessary to give ten liters of carbon dioxide when heated?

Exercise No. 29.

Represent by equations, all the reactions involved in this exercise. Name the alkaline earth metals. Describe the

appearance of calcium hydroxide. How is lime made? How soluble is calcium hydroxide? What is lime water and how is it made? How is sodium hydroxide made? How did you make calcium carbonate? How did you make strontium sulphate? How do calcium compounds color the flame? How could you distinguish between a calcium and a strontium compound? A hydrochloric acid solution contains 20% acid and has a specific gravity of 1.1, how many c.c. of such a solution would be necessary to dissolve ten grams of calcium hydroxide? **155.**

Exercise No. 30.

Represent by equations, all the reactions involved in this exercise. What oxide does magnesium form? Is it basic or acidic? What is formed when it dissolves in water? How many c.c. of a hydrochloric acid solution having a specific gravity of 1.1 and containing 20% acid would be necessary to form ten liters of hydrogen when acting on a metal? **155.** What precipitate is formed when ammonium hydroxide is added to a solution of magnesium chloride? **147.** What is the nature of the action of dilute nitric acid on zinc? What is the precipitate that is formed at first when sodium hydroxide acts on a solution of a zinc salt? Has this precipitate any acid properties? Why does it dissolve when an excess of sodium hydroxide is added? What compound is formed? How could you separate zinc and copper by means of hydrogen sulphide? What is the color of cadmium sulphide? What two nitrates does mercury form and under what circumstances would each result?

Exercise No. 31.

Represent by equations, all the reactions involved in this exercise. What is the nature of the action of dilute nitric acid on copper? How did you make cupric oxide? State three general methods for the preparation of metallic oxides. How did you make cuprous oxide? Why does iron precipitate copper from solution? **121.** What is the most striking reaction of soluble copper compounds? What is the nature of the action of nitric acid on silver? How can you distinguish between

silver and lead chloride? How could you separate silver, lead and mercurous mercury? How many c.c. of hydrochloric acid solution having a specific gravity of 1.1 and containing 20% acid would be necessary to precipitate all the silver in 100 c.c. of a silver nitrate solution having a specific gravity of 1.01 and containing 1% silver nitrate?

Exercise No. 32.

Represent by equations, all the reactions involved in this exercise. What oxides do iron and manganese form? Are these basic or acidic? Does manganese form an acidic oxide? How did you make ferrous sulphate? How else might you have made it? What is green vitriol? Describe ferrous hydroxide. How could you change ferrous sulphate into ferric sulphate? How might you change ferrous chloride to ferric chloride? How change ferric chloride to ferrous chloride? What precipitate is formed when hydrogen sulphide is passed into ferric chloride solution? How would you detect the presence of either ferrous or ferric iron, each in the presence of the other? What is the action of manganese dioxide on sulphurous acid; of potassium permanganate? Give the formulas of ten oxidizing agents and state how they oxidize.

Exercise No. 33.

Represent by equations, all the reactions involved in this exercise. What oxides do aluminum and chromium form? What reagents will dissolve aluminum? Is aluminum very basic? Does it ever act acidic? What is the general nature of an alum? What elements form alums? How do aluminum salts in solution act toward litmus paper? **148.** How could you change basic chromium to acid chromium? Describe the changes that take place when alcohol, potassium dichromate and sulphuric acid are heated together? How many c.c. of a potassium dichromate solution having a specific gravity of 1.03 and containing 5% of the salt would be necessary to oxidize all the iron in 100 c.c. of a ferrous sulphate solution having a specific gravity of 1.05 and containing 10% of green vitriol?

Exercise No. 34.

What metals color the flame red; yellow; blue or green; purple? What are the insoluble chlorides? What are the insoluble sulphates? What insoluble sulphides are white; brown; black; yellow; orange; pink? What oxide is yellow when hot and white when cold? What insoluble hydroxides are white; bluish; brown; green? What substance in solution gives a precipitate of sulphur as well as an odor of sulphur dioxide when treated with an acid?

Exercise No. 35.

What are the general methods for making basic oxides? State seven methods for making salts. Before selecting a method for making a salt, what particular things must be considered? How are insoluble salts made? How are the soluble hydroxides made? How are the insoluble hydroxides made? How could an acid salt be made?

PART II.

FUNDAMENTAL IDEAS.

DEFINITIONS.

1. **Matter** is that which has weight.

2. **An element** is a substance that contains only one kind of matter. There are about seventy-five elements, of which only about twenty-five are common.

Examples of common elements are: iron, copper, lead, tin, gold, and silver. **167-168.**

3. **A symbol** of an element is a sign that stands for that substance. It ordinarily consists of one or two characteristic letters of the English or Latin name.

4. Symbols and names of some common elements, in groups. **167.**

Group 1.	Group 2.	Group 3.	Group 4.
H, Hydrogen.	Mg, Magnesium.	B, Boron.	C, Carbon.
Na, Sodium.	Ca, Calcium.	Al, Aluminum.	Si, Silicon.
K, Potassium.	Zn, Zinc.		Sn, Tin.
NH_4^*, Ammonium.	Ba, Barium.		Pb, Lead.
Ag, Silver.	Hg, Mercury.		

Group 5.	Group 6.	Group 7.	Group 8.
N, Nitrogen.	O, Oxygen.	F, Fluorine.	Fe, Iron.
P, Phosphorus.	S, Sulphur.	Cl, Chlorine.	Ni, Nickel.
As, Arsenic.	Cr, Chromium.	Br, Bromine.	Co, Cobalt.
Sb, Antimony.		I, Iodine.	
Bi, Bismuth.			

* Not an element, but a compound of N and H that acts like one.

5. **Properties of matter.** Those qualities which are peculiar to it and characteristic of it. An exact description of any kind of matter involves telling what it looks like, smells like, tastes like, how it behaves at different temperatures and when mixed with other substances.

6. All matter is supposed to be made up of a very large number of very small particles called molecules, and these molecules to be made up of still smaller particles called atoms.

7. **A molecule** is the smallest particle of matter that can exist and still retain all the properties of that kind of matter. If the substance is an element, the molecule will consist of one or more atoms of that element; if the substance is a chemical compound, the molecule will contain at least one atom of every element contained in the substance.

8. **An atom** is the smallest particle of an element which forms part of a molecule.

9. **The chemical composition** of a substance means the relative proportions, by weight, of all the elements contained in it.

10. **A physical change** is one in which the chemical composition of the body is not altered. Thus, a piece of matter may be set in motion, be heated, electrified, magnetized, melted or dissolved without in any way changing its chemical composition. Such changes are physical.

11. **A chemical change** is one in which the chemical composition of the body is altered. Such changes are always accompanied by some change in the total energy which is associated with the reacting substances. That is, there is an evolution or absorption of heat, light or electricity. For example; if sugar be heated hot enough, it changes into carbon and water. Or if carbon be heated in the air, it changes to a colorless gas, called carbon dioxide. The rusting of iron, the souring of milk, and the fermenting of cider are also examples of chemical changes.

12. If two or more elements be mixed in such a way that they cannot be separated again by mechanical means, a **chemical compound** is said to be formed. For example; if sulphur and iron be heated together, a compound called iron sulphide is formed; and no mechanical process can again separate them. Such a compound contains only one kind of molecule.

13. **A mechanical mixture** is one containing the elements mixed in such a way that they may be separated by mechanical means. The iron and sulphur could readily have been separated by means of a magnet before they had been heated. Most mechanical mixtures are mixtures of chemical compounds. When viewed under the microscope such mixtures show different kinds of particles. In the case of solutions there is considerable difficulty in distinguishing them from chemical compounds, for all parts of the same solution have exactly the same chemical composition and it is impossible to distinguish different kinds of articles by means of a microscope. However, the solvent can usually be entirely evaporated in the pure condition, leaving the other material by itself, either in the liquid or solid condition. A mechanical mixture contains two or more different kinds of molecules.

14. **An atomic weight** is a number expressing the ratio between the weight of the smallest particle of an element forming part of a molecule and the weight of an atom of hydrogen. In other words, it is the ratio of the amount of matter in an atom to the amount of matter in one atom of hydrogen. Sometimes one-sixteenth the weight of the oxygen atom is taken as the unit, in place of the hydrogen atom.

15. **A molecular weight** is a number which shows how many times heavier a molecule of a substance is than one atom of hydrogen. It is equal to the sum of the atomic weights of all the atoms in the molecule.

FULL MEANING OF SYMBOLS AND FORMULAS.

16. **Full meaning of the symbol of an element.**

1. It is a sign that stands for that substance.
2. It is a sign that stands for one atom of that substance.
3. It is a sign that stands for a certain weight of that substance.

For example: S is a sign that stands for sulphur, for one atom of sulphur, and for 32 weights of sulphur.

The weight corresponding to a given symbol is the same as the atomic weight of the element, and may be found in any table of elements. **168.**

17. **The formula** of a compound is a sign that stands for one molecule of that substance. It is made up of the symbols of the elements contained in the compound, with a number below and at the right of each symbol, indicating how many atoms of that element are contained in the molecule.

18. **Formulas of some common compounds.**

H_2O, Hydrogen oxide or water.
CO_2, Carbon dioxide or carbonic acid gas.
MgO, Magnesium oxide or magnesia.
NaCl, Sodium chloride or common salt.
HCl, Hydrogen chloride, hydrochloric acid or muriatic acid.
HNO_3, Hydrogen nitrate or nitric acid.
H_2SO_4, Hydrogen sulphate, sulphuric acid or oil of vitriol.

The Full Meaning of the Formula of a Compound.

19. This can best be shown by giving the exact meaning of some formula.

1. H_2SO_4 is a sign that stands for sulphuric acid.
2. It stands for one molecule of sulphuric acid.
3. It stands for 98 weights of sulphuric acid.

It states that,

4. Sulphuric acid is made of hydrogen, sulphur and oxygen.
5. That one molecule of sulphuric acid contains two atoms of hydrogen, one of sulphur and four of oxygen.

6. That these elements are present in the proportion of 2 weights of hydrogen, 32 weights of sulphur and 64 weights of oxygen.

The student should practice giving the full meaning of other formulas.

20. **Analysis** is the process of finding out the relative proportions, by weight, of the elements contained in a substance.

The results of an analysis are always expressed in per cents.

21. **To change the Composition of a Substance as shown by a Formula to a percentage Basis.**

Example. Given the formula of sulphuric acid, H_2SO_4, to find its percentage composition.

$$\begin{aligned} 2H &= 2 \\ S &= 32 \\ 4O &= 64 \\ & \overline{98} \end{aligned}$$

There are 98 weights in the molecule altogether; of these 2 are hydrogen; therefore $\frac{2}{98}$ reduced to a decimal and multiplied by 100 must equal the percentage of this constituent.

Proceeding in the same way with the other elements we have,

$$\begin{aligned} H &= \frac{2}{98} = .0204 \text{ or } 2.04\% \\ S &= \frac{32}{98} = .3265 \text{ or } 32.65\% \\ O &= \frac{64}{98} = .6531 \text{ or } 65.31\% \\ & \overline{100.00} \end{aligned}$$

Exercise.

22. Compute the percentage composition of substances from their formulas.

Derivation of simple Formulas from Analysis.

23. If the molecule happens to contain a single atom of any one kind, it is a simple matter to derive the formula from the analysis. It is only necessary to divide each given percentage

by the corresponding atomic weight and the results will be the relative numbers of atoms in the molecule. Each number will be a simple multiple of the smallest, and the actual numbers of atoms is readily seen by inspection. For the treatment of cases involving more complex molecules, see **95.**

Exercise.

24. Derive the formulas corresponding to the following analyses:

Hydrogen, 2.74%, chlorine, 97.26%; hydrogen, 11.11%, oxygen, 88.89%; hydrogen, 17.63%, nitrogen, 82.35%; hydrogen, 25.00%, carbon, 75.00%.

25. Metals and Non-metals. 168.

All the elements may be divided into two classes; metals and non-metals. In addition to the common elements like copper, lead, tin, iron, and silver, the student may remember that mercury, bismuth and antimony are metals. Also that most of the elements whose names end in "um" are metals. **168.**

26. A **binary compound** is one containing only two elements.

In naming a binary compound the name of the metal is placed first, unchanged, followed by the name of the non-metal with its name changed to an ending in "ide," e.g., a compound of sodium and chlorine is called sodium chloride; or of iron and oxygen, iron oxide, etc. While almost any element may form a binary compound with any other element, by far the most common compounds are those of the metals with the non-metals.

Where both of the elements are non-metals, the name of either may be placed first, followed by the name of the other with an ending in "ide," e.g., hydrogen nitride or nitrogen hydride.

27. The Law of Definite Proportions.

If equal weights of two elements be caused to combine it will be found that one of them will entirely enter into combination and some of the other will be left uncombined. No matter how

the quantities be varied one of the substances will always entirely enter into combination. If the resulting compound be analysed, the constituent elements will always be found to be present in definite proportions by weight. If the two original elements be combined in these proportions they will exactly unite to form the compound. Thus if four weights of sulphur and seven of iron be heated together they exactly combine to form a compound known as iron sulphide.

In a given chemical compound the elements are always present in the same fixed proportions by weight.

28. The Law of Multiple Proportions.

Analysis of all the known binary compounds shows that there are many cases of the same two elements uniting in different proportions. In such cases it is found that when a fixed quantity of one element unites with variable quantities of another, the weights of the variable element are simple multiples of the smallest. Thus, seven weights of nitrogen combine with four, eight, twelve and sixteen weights of oxygen, forming four different oxides.

29. The Atomic Hypothesis of Dalton.

As early as the fifth century B. C. certain philosophers considered all matter to be made up of small particles called atoms. These atoms were all assumed to be of the same kind and the same size. Different kinds of matter were supposed to result from the atoms approaching or receding from each other.

In 1807 Dalton proposed the atomic theory which is in use today and is the basis of all modern theoretical chemistry. According to his theory, all the atoms of a given element are of the same size and the same weight but atoms of different elements have different weights. If two or more atoms of the same kind unite, they constitute a molecule of an element. If two or more atoms of different kinds unite, they form a molecule of a compound.

This theory agrees with the **law of definite proportions** and the **law of multiple proportions**, for if two atoms unite to form a molecule of a compound, the proportions of the elements in a mass of the substance would be identical with the proportions

by weight of the elements in the molecule. Again, if a fixed weight of one element unites with variable weights of another element to form several different compounds, the ratios of the weights of the variable element to each other are of the nature of simple numbers.

This is well illustrated by the oxides of nitrogen. 28 weights of nitrogen combine with 16, 32, 48, 64 and 80 weights of oxygen. The ratio of the last four weights to the first is in every case a simple whole number.

COMBINING POWER OR VALENCE.

30. It is possible to determine exactly the molecular weight of a compound by several processes. **92-93.** This, taken in connection with the analysis and the atomic weights, makes it possible to arrive at the correct formula. **95.**

The formulas of all the binary compounds have been determined; and from an inspection of them it appears, that the atoms of the various elements have different powers of holding each other in combination in the molecules.

31. **Combining power** or **Valence** is the power that an atom has of combining with other atoms to form molecules. All the elements in the same group have the same combining power. The valence of an element is measured by the number of atoms of hydrogen, or their equivalent, which its atom can hold in combination. Compounds with oxygen or hydrogen are usually taken to show the valence, although any element whose valence is known will serve just as well.

Combining power of the elements by groups.

The elements in group one have a power of one.

The elements in group two have a power of two.

The elements in group three have a power of three.

The elements in group four have a power of four.

The elements in group five have powers of three and five.

The elements in group six have powers of two, four and six, except oxygen which has only two.

The elements in group seven have powers of one, three, five and seven.

The elements in group eight (see **167**) have powers of two and three.

In cases where an element has several combining powers, the one to be used in the following exercises, unless stated to the contrary, is the lowest.

DERIVATION OF FORMULAS OF BINARY COMPOUNDS FROM THE KNOWN COMBINING POWERS OF THE ELEMENTS.

32. **Almost any two elements can combine to form a binary compound. The total valence or combining power of the atoms of any one kind forming a molecule of a binary**

compound is the least common multiple of the two combining powers concerned.

In other words, there must be enough atoms of the first kind present to give a total valence equal to that of one or more atoms of the second kind. If this valence is the least possible one that gives equivalent combining powers, the resulting formula is the **simplest one possible.**

The simplest formula is not necessarily the correct one. The true formula may be a multiple of the simplest. **95.**

Sodium has a power of one, being in the first group, and it will therefore exactly combine with one atom of chlorine, which is in the seventh group, but has a low common power of one. The formula would be NaCl.

Magnesium, with a power of two, will combine with two chlorine atoms, giving the formula $MgCl_2$.

Aluminum, in the third group, with a power of three, would combine with three chlorine atoms, giving the formula $AlCl_3$.

The formula of a compound of carbon and chlorine would be CCl_4.

The formula of a compound of aluminum and oxygen would be Al_2O_3, for the least common multiple of the two combining powers would be six, and it takes two aluminum atoms to give an equivalent power. Oxygen, in the sixth group, always has a combining power of two and it will take three atoms to give a valence of six.

The sub-numbers in the formula are ordinarily the same as the combining powers of the two elements, reversed. If one is divisible by the other, the simplest ratio is taken in the reversed order, e.g., carbon has a power of four and oxygen of two; the formula is not C_2O_4, but CO_2.

When two elements form several different binary compounds, it is customary to distinguish between them by placing the prefixes, mon-; di-; tri-; tetra-; penta-; etc., before the name of the variable element. Thus SO_2 is sulphur dioxide, SO_3 is sulphur trioxide, P_2O_5 is phosphorus pentoxide, P_2O_3 is phosphorus trioxide.

Remember always, *that formulas are not originally derived from valence but valence from formulas. And that the theory of valence is only taught in order to relieve the student of the*

necessity of remembering the combining proportions of the elements in thousands of well known compounds.

Exercise.

33. Give the formulas and names of all the binary compounds of the metals with the non-metals.

CHEMICAL EQUIVALENCE.

34. A quantity of one element is said to be equivalent to a certain quantity of another element when there are enough atoms of each present to give the same combining power.

One atom of sodium is equivalent to one atom of potassium.

Two atoms of sodium are equivalent to one atom of magnesium.

Two atoms of magnesium are equivalent to one atom of carbon.

Three atoms of zinc are equivalent to two atoms of aluminum.

The term "atom" is used in the sense of a certain number of weights. Thus, three atoms of zinc means three times 65 weights of zinc; and two atoms of aluminum means two times 27 weights of aluminum.

The Number of Grams of any one Element equivalent to a Certain Number of Grams of any other Element.

How many grams of aluminum are equivalent to ten grams of zinc?

Two atoms of aluminum are equivalent to three atoms of zinc. That is, 54 weights of aluminum are equivalent to 195 weights of zinc. One weight would therefore be equivalent to $\frac{1}{54}$ as much, or $\frac{195}{54}$, and as many would be equivalent to ten as $\frac{195}{54}$ is contained in ten, or $10 \times \frac{54}{195}$.

Or, since 195 weights of zinc are equivalent to 54 weights of aluminum, one weight would be equivalent to $\frac{1}{195}$ as much or $\frac{1}{195} \times 54$, and ten weights would be equivalent to $10 \times \frac{54}{195}$.

Exercise.

35. Give the chemical equivalence between every two elements, and give the expression for the answer when ten grams of either element is given. That is, ten grams of sodium is equivalent to how many grams of aluminum, etc.?

36. **A reaction** is a chemical change.

37. **A reagent** is any substance that takes part in a reaction.

38. **A chemical equation** is one, on the left hand side of which are placed the symbols of the molecules of the substances used, and on the right hand side are placed the symbols of the molecules of the substances formed. A number is placed before each symbol indicating the least possible number of *molecules* of that substance which could take part in that reaction.

An equation can never be a complete description of a reaction for it only tells what substances take part in the reaction, and their proportions. Besides these facts, a complete description must state the circumstances under which the particular reaction under consideration takes place, especially with regard to the temperature, the pressure and the concentrations of the various substances concerned.

According to this definition, it is necessary to know how many atoms there are in the molecules of the elements. As there is considerable uncertainty as to the number of atoms in the molecules of many of the elements, for the present it will be stated that the elementary gases, oxygen, hydrogen, nitrogen, chlorine and several other elements (see **168**) have two atoms in the molecule, and when these elements appear *by themselves* in an equation they should be written, O_2, H_2, N_2 and Cl_2. On account of the uncertainty it is preferred to write all other elements as atoms.

When all the facts regarding a chemical reaction are known it is possible to represent that reaction by a chemical equation.

For example, 46 weights of sodium exactly unite with 16 weights of oxygen to form 62 weights of sodium oxide. Since the atomic weight of sodium is 23, it will take two atoms of sodium and one atom of oxygen to form one molecule of sodium oxide.

That could be expressed:

$2Na + O = Na_2O$; but since we are going to express oxygen in the form of a molecule we shall double the quantities and write it, $4Na + O_2 = 2Na_2O$.

Again, to be more general, suppose that actual experiment shows that 26.47 grams of aluminum exactly combine with 23.53 grams of oxygen to form 50.00 grams of aluminum oxide, and that it is desired to express this reaction by an equation. Dividing the first two numbers by the corresponding atomic weights of the elements and the last by the molecular weight of aluminum oxide, we obtain three numbers which show at once the relative numbers of atoms of the elements compared to the number of molecules of the compound. These numbers are,

Al	O	Al_2O_3
.9804	1.4700	.4902

This means that two atoms of aluminum unite with three of oxygen to form one molecule of aluminum oxide.

Since we desire to write a molecular equation as far as possible we double all these quantities and write the equation,

$$4Al + 3O_2 = 2Al_2O_3.$$

EQUATIONS REPRESENTING THE FORMATION OF BINARY COMPOUNDS.

39. Since the combining power of the various elements has been well established by experiment in the past, it is now possible to predict their formation exactly and represent the corresponding reactions by equations.

Nitrogen unites with oxygen. Two nitrogen atoms are chemically equivalent to three oxygen atoms, and unite to form one molecule of nitrogen oxide. Thus:

$2N + 3O = N_2O_3$, but in this equation we have not represented oxygen and nitrogen as molecules; we therefore rewrite the equation, doubling the quantities; so,

$$2N_2 + 3O_2 = 2N_2O_3.$$

Exercise.

40. Write equations showing the formation of the binary compounds resulting from the combinations of the metals with

the non-metals. Especially show the formation of the oxides, sulphides, and chlorides.

$$2Na + Cl_2 = 2NaCl.$$
$$4Al + 3O_2 = 2Al_2O_3.$$
$$C + O_2 = CO_2.$$
$$Zn + S = ZnS.$$

Etc.

FULL MEANING OF A CHEMICAL EQUATION.

41. The equation,

$$4Al + 3O_2 = 2Al_2O_3 \text{ means}$$

1. That aluminum combines with oxygen to form aluminum oxide.

2. That four atoms of aluminum exactly unite with three molecules of oxygen to form two molecules of aluminum oxide.

3. In other words, 108 weights of aluminum exactly unite with 96 weights of oxygen to form 204 weights of aluminum oxide.

4. Therefore that ten grams of aluminum would exactly unite with $\frac{96}{108} \times 10$ grams of oxygen and form $\frac{204}{108} \times 10$ grams of aluminum oxide.

Also that 10 grams of oxygen would combine with $\frac{108}{96} \times 10$ grams of aluminum and form $\frac{204}{96} \times 10$ grams of aluminum oxide and finally that 10 grams of aluminum oxide would be formed by the combination of $\frac{108}{204} \times 10$ grams of aluminum with $\frac{96}{204} \times 10$ grams of oxygen.

42. Give the full meaning of every equation you have written, working it out for ten grams as above.

FORMULAS OF COMPOUNDS CONTAINING MORE THAN TWO ELEMENTS.

43. A compound containing more than two elements is called a **Ternary** compound.

44. It is found that when the electric current is passed through many fused compounds or compounds in solution,

that they are broken up into two parts; one part, consisting of the metal or hydrogen, going to the negative electrode, and all the rest of the substance going to the positive electrode. **124.**

45. **A radical** is an atom or a group of atoms which enters various reactions as a whole. Ex.—H; Na; NH_4; Cl; SO_4; NO_3.

46. **A basic radical** is one that will move toward the negative electrode, when an electric current is passed through a solution of the substance. Hydrogen and nearly all the metals constitute the basic radicals.

47. **An acid radical** is one that moves toward the positive electrode, when an electric current is passed through a solution of the substance. Nearly all of the acid radicals contain several atoms.

48. Radicals may be considered to have combining powers exactly like the elements themselves, and to unite with each other to form compounds in a similar way.

49. **Names of radicals.** The names of the radicals given in the following list are the names which, put together just as they are, will give the proper chemical name of the compound which any two radicals form.

50. In order that the student may become familiar at once with many of the names and formulas of compounds that he must necessarily make use of in the laboratory work, the following list of radicals should be committed to memory. Extensive drill should also be given in combining the basic with the acid radicals and in naming the compounds so indicated. The derivation of these and other radicals will be given later.

LIST OF RADICALS WITH NAMES.

Basic Radicals.		Acid Radicals.	
Group I. Power of One.		Group I. Power of One.	
H,	Hydrogen.	$C_2H_3O_2$,	Acetate.
Na,	Sodium.	Br,	Bromide.
K,	Potassium.	ClO_3,	Chlorate.
NH_4,	Ammonium.	Cl,	Chloride.
Ag,	Silver.	OH,	Hydroxide.
Hg,	Mercurous.	I,	Iodide.
Cu,	Cuprous.	NO_3,	Nitrate.

Basic Radicals.	Acid Radicals.
Group 2. Power of Two.	**Group 2. Power of Two.**
Mg, Magnesium.	CO_3, Carbonate.
Ca, Calcium.	O, Oxide.
Zn, Zinc.	SO_4, Sulphate.
Ba, Barium.	S, Sulphide.
Hg, Mercuric.	
Cu, Cupric.	
Fe, Ferrous.	
Pb, Lead.	
Group 3. Power of Three.	**Group 3. Power of Three.**
Al, Aluminum.	PO_4, Ortho-phosphate.
Fe, Ferric.	$Fe(CN)_6$, Ferricyanide.
	Group 4. Power of Four.
	$Fe(CN)_6$, Ferrocyanide.

51. If a radical has two different combining powers, it is evident that it will form two classes of compounds; and therefore it is necessary that two different names be assigned to it, in order to signify which compound is being considered. Thus iron forms two oxides; ferrous oxide, with the formula FeO, and ferric oxide, with the formula Fe_2O_3.

COMBINATIONS OF RADICALS.

52. Basic radicals unite with acid radicals.

The total combining power of the basic radical must be exactly equal to the total combining power of the acid radical, exactly as in the formation of a binary compound.

When it is necessary to represent a compound radical as occurring several times in a formula, the whole radical is put in parenthesis with a number below and at the right, showing how many times it is found.

The name is formed by simply putting together the two names of the radicals, placing that of the basic radical first.

Exercise.

53. Give the names and formulas of all the radicals. Give the names and formulas of the radicals by groups.

Exercise.

54. Give the formulas and names of all the compounds of the basic radicals with the acid radicals.

To illustrate, take one typical radical from each group.

$H.'$	$Cl.'$
$Na.'$	$NO_3.'$
$Ba.''$	$SO_4.''$
$Al.'''$	$PO_4.'''$
	$Fe(CN)_6.''''$

The primes indicate the combining power to be used in this particular case.

HCl, Hydrogen chloride.
HNO_3, Hydrogen nitrate.
H_2SO_4, Hydrogen sulphate.
H_3PO_4, Hydrogen phosphate.
$H_4Fe(CN)_6$, Hydrogen ferrocyanide.

In this set it takes a number of H's equal to the combining power of the acid radical to form the molecule.

$NaCl$, Sodium chloride.
$NaNO_3$, Sodium nitrate.
Na_2SO_4, Sodium sulphate.
Na_3PO_4, Sodium phosphate.
$Na_4Fe(CN)_6$, Sodium ferrocyanide.

Here again, since the combining power of sodium is one, it takes a number of sodium radicals equal to the combining power of the acid radical to form the molecule.

$BaCl_2$, Barium chloride.
$Ba(NO_3)_2$, Barium nitrate.
$BaSO_4$, Barium sulphate.
$Ba_3(PO_4)_2$, Barium phosphate.
$Ba_2Fe(CN)_6$, Barium ferrocyanide.

Barium has a power of two, chloride has a power of one; two chloride radicals will be equivalent to one barium radical

and will exactly unite with it to form a molecule. The same with the nitrate. Sulphate has a power of two and is equivalent to the barium radical at once. The phosphate radical has a power of three. The least common multiple of the two powers concerned, two and three, is six. It will take three bariums to equal six; and two phosphates.

$AlCl_3$, Aluminum chloride.
$Al(NO_3)_3$, Aluminum nitrate.
$Al_2(SO_4)_3$, Aluminum sulphate.
$AlPO_4$, Aluminum phosphate.
$Al_4(Fe(CN)_6)_3$, Aluminum ferrocyanide.

In the last case, the two combining powers concerned are three and four; the least common multiple of these two is twelve; three goes in twelve four times; there are therefore four aluminums to be used; four goes in twelve three times, which calls for three ferrocyanide radicals.

55. **An acid** is a substance that is generally sour to the taste, turns blue litmus paper red and neutralizes a base to form water and a salt. It always contains hydrogen combined with an acid radical.

56. **A base** is a compound containing the radical OH, and neutralizes an acid forming water and a salt. It *generally* consists of a basic radical joined to OH.

57. **An alkali** is a strong caustic base which is soapy to the taste and feeling and will turn red litmus blue. The common alkalis are sodium, potassium, and ammonium hydroxides.

58. **A salt** is a compound of any basic radical except H, with any acid radical except O or OH, and is always formed together with water when an acid combines with a base or a basic oxide.

Exercise.

59. Give the names of the compounds of the basic radicals with the acid, stating in each case whether the compound is an acid, a base, or a salt.

60. **A basic oxide** is one that unites with an acid to form water and a salt. Most of the oxides of the metals are basic oxides.

MgO; ZnO; HgO; PbO; CuO.

61. **An acidic oxide** is one that unites with a base to form water and a salt. Most of the oxides of the non-metals are acidic oxides.

$$CO_2;\ SiO_2;\ SO_2;\ SO_3;\ P_2O_5.$$

62. **An alkaline oxide** is a basic oxide that unites with water to form an alkali.

$$Na_2O;\ K_2O;\ CaO;\ BaO;\ MgO.$$

63. **An acid anhydride** is an acidic oxide that unites with water to form an acid.

$$CO_2;\ SO_2;\ SO_3;\ P_2O_3;\ P_2O_5.$$

REACTIONS RESULTING IN THE FORMATION OF SALTS.

64a. THE ACTION OF A METAL ON AN ACID.

In this reaction the metal takes the place of an equivalent quantity of hydrogen in the acid, and the hydrogen is set free. In case there is an oxidizing agent present, the hydrogen would be oxidized to water and would not appear in the free condition.

$$2Na + 2HCl = 2NaCl + H_2.$$
$$Zn + 2HCl = ZnCl_2 + H_2.$$
$$Zn + H_2SO_4 = ZnSO_4 + H_2.$$
$$2Al + 3H_2SO_4 = Al_2(SO_4)_3 + 3H_2.$$

Combine each of the following basic radicals with each of the acid radicals, and write equations indicating the formation of the resulting salts, by the above method.

Basic radicals.	Acid radicals.
$Na.'$	$Cl.'$
$K.'$	$NO_3.'$
$Ba.''$	$SO_4.''$
$Mg.''$	$PO_4.'''$
$Al.'''$	$Fe(CN)_6.''''$

b. THE ACTION OF A BASIC OXIDE ON AN ACID.

In this case the oxygen of the oxide unites with the hydrogen of the acid, forming water, thus leaving the basic radical of the oxide free to combine with the acid radical of the acid.

$$Na_2O + 2HCl = 2NaCl + H_2O.$$
$$ZnO + 2HCl = ZnCl_2 + H_2O.$$
$$Al_2O_3 + 3H_2SO_4 = Al_2(SO_4)_3 + 3H_2O.$$

Write equations showing the formation of the same twenty-five salts by this method.

c. THE ACTION OF A BASE ON AN ACID.

In this case, the hydroxide of the base unites with the hydrogen of the acid, forming water, thereby leaving the basic radical of the base free to combine with the acid radical of the acid.

$$NaOH + HCl = NaCl + HOH.$$
$$Zn(OH)_2 + 2HCl = ZnCl_2 + 2HOH.$$
$$2Al(OH)_3 + 3H_2SO_4 = Al_2(SO_4)_3 + 6HOH.$$

Write equations showing the formation of the same twenty-five salts by this method.

d. THE ACTION OF AN ACID ON THE SALT OF A MORE VOLATILE ACID.

In this reaction the hydrogen of the acid unites with the acid radical of the salt and forms another acid, which is either unstable and breaks up into volatile gaseous compounds at once, or evaporates as an acid either with or without the application of heat, leaving the acid radical of the original acid combined with the basic radical of the salt.

Carbonic acid, or hydrogen carbonate, H_2CO_3, is one of the most unstable and volatile acids, therefore almost any acid will decompose a carbonate. Sulphuric and phosphoric acids are very stable and non-volatile. Therefore, when almost any salt is treated with sulphuric acid, another acid is set free and a sulphate is formed.

$$Na_2CO_3 + 2HCl = 2NaCl + \overbrace{H_2O + CO_2}.$$
$$ZnCO_3 + 2HCl = ZnCl_2 + \overbrace{H_2O + CO_2}.$$
$$2Al(NO_3)_3 + 3H_2SO_4 = Al_2(SO_4)_3 + 6HNO_3.$$

Write equations showing the formation of the same twenty-five salts by this method, making use of the carbonates containing the required basic radical.

In order that the student may intelligently make use of this kind of reaction, it is important that he should have some idea of the relative degrees of volatility or stability of some of the

more common acids. For this purpose all acids might be divided into three classes, those that are extremely non-volatile and stable, those that are easily volatile but of sufficient stability to admit of being prepared in the free condition, and those that are so unstable that they decompose, the moment they are formed, into water and oxides of non-metals.

Examples of acids that are non-volatile and stable.

H_2SO_4, Hydrogen sulphate or sulphuric acid.
H_3PO_4, Hydrogen ortho-phosphate or ortho-phosphoric acid.

Examples of acids that are volatile and stable.

HNO_3, Hydrogen nitrate or nitric acid.
HCl, Hydrogen chloride or hydrochloric acid.
$HC_2H_3O_2$, Hydrogen acetate or acetic acid.

Examples of acids that are unstable.

H_2CO_3, Hydrogen carbonate or carbonic acid.
H_2SO_3, Hydrogen sulphite or sulphurous acid.
HNO_2, Hydrogen nitrite or nitrous acid.

e. THE MIXING OF TWO SOLUTIONS, ONE OF WHICH CONTAINS THE BASIC RADICAL AND THE OTHER THE ACID RADICAL OF AN INSOLUBLE COMPOUND.

In this case the insoluble compound separates as a precipitate at once, and may be filtered off. If the two substances have been mixed in chemically equivalent proportions, the filtrate may be evaporated and the other compound obtained. The insoluble compound may be an acid, a base, or a salt. In making use of this method, it is necessary to know whether or not a substance is soluble. For this purpose a table of solubilities may best be consulted. **165.**

Silver chloride, barium sulphate, and calcium carbonate are almost completely insoluble in water; some one of them may be used in preparing any salt by this method. It is also well to remember that

ALL THE HYDROGEN, SODIUM, POTASSIUM AND AMMONIUM COMPOUNDS ARE SOLUBLE, ALSO THE CHLORATES, ACETATES AND NITRATES AND ALL THE CHLORIDES EXCEPT THOSE OF SILVER, LEAD, AND MERCUROUS MERCURY.

In writing equations showing the formation of insoluble compounds, the latter should be indicated by a line drawn underneath the formula: e.g. It is desired to show the formation of sodium chloride by this method.

Making use of the insoluble compound, barium sulphate, we see that we must have two soluble compounds which by their combination will form sodium chloride and barium sulphate; these two compounds must therefore be barium chloride and sodium sulphate. We know these are both soluble in water, because, according to the rule, one is a chloride and the other a sodium compound. We therefore have

$$Na_2SO_4 + BaCl_2 = 2NaCl + \underline{BaSO_4}.$$
$$ZnSO_4 + BaCl_2 = ZnCl_2 + \underline{BaSO_4}.$$
$$2AlCl_3 + 3Ag_2SO_4 = Al_2(SO_4)_3 + \underline{6AgCl}.$$

Show, by equations, the formation of the same twenty-five salts, using in each case one of the above three insoluble compounds.

f. THE DIRECT UNION OF A BASIC OXIDE WITH AN ACIDIC OXIDE.

This method applies only to the formation of oxygen salts.

$$Na_2O + SO_3 = Na_2SO_4.$$
$$Al_2O_3 + 3N_2O_5 = 2Al(NO_3)_3.$$
$$PbO + SiO_2 = PbSiO_3.$$

Combine each of the following acidic oxides with each of the basic oxides.

Acidic oxides.		Basic oxides.
N_2O_5	forms nitrates.	Na_2O.
P_2O_5	forms phosphates.	K_2O.
As_2O_3	forms arsenites.	BaO.
SO_3	forms sulphates.	Al_2O_3.

Acidic oxides.		Basic oxides.
CrO_3	forms chromates.	PbO.
SiO_2	forms silicates.	Fe_2O_3.

New radicals. AsO_3''', Arsenite. CrO_4'', Chromate. SiO_4'''', Orthosilicate. NO_2, Nitrite. SO_3, Sulphite.

g. THE ACTION OF AN ACID ANHYDRIDE ON A BASE.

In this case water and a salt are formed. The action is essentially the same as the action of an acid on a base, for the acid anhydride acting on the water at once forms the corresponding acid.

Write equations showing the combination of the following acid anhydrides with the indicated bases:

Acid anhydrides.	Bases.
CO_2.	NH_4OH.
SiO_2.	$NaOH$.
N_2O_3.	KOH.
SO_2.	$Ca(OH)_2$.
SO_3.	$Ba(OH)_2$.
CrO_3.	$Sr(OH)_2$.

In writing equations illustrating this method, it will be best for the student at first to write an equation showing the formation of the corresponding acid by the union of the anhydride with water, then write a second equation showing the combination of the base with the quantity of acid formed. Adding these two equations algebraically gives the required final question.

Thus,

$$N_2O_3 + H_2O = 2HNO_2.$$
$$2HNO_2 + 2NaOH = 2NaNO_2 + 2H_2O.$$

$$N_2O_3 + 2NaOH = 2NaNO_2 + H_2O.$$

or,

$$P_2O_5 + 3H_2O = 2H_3PO_4.$$
$$2H_3PO_4 + 3Ca(OH)_2 = Ca_3(PO_4)_2 + 6H_2O.$$

$$P_2O_5 + 3Ca(OH)_2 = Ca_3(PO_4)_2 + 3H_2O.$$

It ought not to be necessary to write the two preliminary equations,

$$SO_3 + Ca(OH)_2 = CaSO_4 + H_2O.$$

OXIDATION.

65. Most of the elements combine directly or indirectly with oxygen to form oxides, and many of them combine with different quantities of oxygen, forming different oxides. In such cases we recognize what may be called the *stage of oxidation*. An element may exist in its highest, lowest or some intermediate stage of oxidation.

Oxidation is the process of adding oxygen to or extracting hydrogen from a substance. It is also the process of changing an element from a lower to a higher stage of oxidation. Thus we might consider chlorine in the compound HCl to be in its lowest stage of oxidation. If this compound is oxidized chlorine is set free. Free chlorine would then represent a higher stage of oxidation than when combined with hydrogen.

66. **An oxidizing agent** is any substance that directly or indirectly brings about oxidation.

67. **Reduction** is the process of extracting oxygen or adding hydrogen to a substance. It is also the process of changing an element from a higher to a lower stage of oxidation.

All oxidizing processes may be divided into two classes, oxidation in the *dry way* and oxidation in the *wet way*.

Oxidation in the Dry Way.

68. In cases of oxidation in the dry way the oxidizing agent is commonly free gaseous oxygen or some oxygen compound which readily gives up some or all of its oxygen. Such reactions take place without the intervention of a liquid and commonly therefore result in temperatures sufficiently high to produce light as well as heat.

Many elements when heated in oxygen or in air combine directly with oxygen, forming oxides. If an excess of oxygen is present the highest stage of oxidation will ordinarily be attained; if an excess of the element is present a lower form of oxide may be produced. Some elements unite so energetically with oxygen that the oxidation proceeds by itself with the production of light as well as heat. Such substances are said to burn. Magnesium, sulphur and carbon burn in air. Chemi-

cal combination accompanied by light and heat is called *combustion.* Burning is simply a common form of combustion.

If a substance consisting of several combustible elements is ignited in air it generally burns, and the products of the burning are the oxides of the elements it contained. Wood, illuminating gas, oil and alcohol are examples of such substances.

Metallic oxides themselves are oxidizing agents if heated with some material that combines very readily with oxygen. Examples of materials that have great power of uniting with oxygen are hydrogen, carbon, sodium, potassium, aluminum and magnesium.

$$2PbO + C = 2Pb + CO_2.$$
$$2Fe_2O_3 + 3C = 4Fe + 3CO_2.$$
$$CuO + H_2 = Cu + H_2O.$$

It should be noted that oxidation is always accompanied by reduction. The oxidizing agent oxidizes the reducing agent and the reducing agent reduces the oxidizing agent.

Some Oxidizing Agents for the Dry Way:

Oxygen.
Air.
$KClO_3$, Potassium chlorate.
KNO_3, Potassium nitrate.
Pb_3O_4, Red lead.
MnO_2, Manganese dioxide.

Equations illustrating oxidation in the dry way, where oxygen is the oxidizing agent.

$$C + O_2 = CO_2.$$
$$2Cu + O_2 = 2CuO.$$
$$4Cu + O_2 = 2Cu_2O.$$
$$2Mg + O_2 = 2MgO.$$
$$2C_2H_6 + 7O_2 = 4CO_2 + 6H_2O.$$
$$C_2H_5OH + 3O_2 = 2CO_2 + 3H_2O.$$

Exercise.

Write equations illustrating the oxidation or burning of the following substances:

Sodium, to form sodium oxide.
Magnesium, to form magnesium oxide.

Zinc, to form zinc oxide.
Aluminum, to form aluminum oxide.
Carbon, to form CO_2.
Nitrogen, to form N_2O_3.
Nitrogen, to form N_2O_5.
Phosphorus, to form P_2O_3.
Phosphorus, to form P_2O_5.
Sulphur, to form SO_2.
Sulphur, to form SO_3.

69. The Oxides of the Elements in Groups in Relation to Acids and Bases.

All the elements may be divided into two classes, metals and non-metals. In general, the oxides of the metals unite directly or indirectly with water to form bases. The oxides of the non-metals unite with water either directly or indirectly to form acids.

The First Group.

All the elements of this group are metals and all form oxides of the general formula R_2O. Copper, in addition to the oxide Cu_2O, forms the oxide CuO. These oxides are all base producing.

$$Na_2O + H_2O = 2NaOH.$$

The Second Group.

All the elements of this group are metals and form oxides of the general formula RO. Mercury, in addition to the oxide HgO, forms the oxide Hg_2O. These oxides are also all base producing.

$$CaO + H_2O = Ca(OH)_2.$$

The Third Group.

Two elements only are considered in this group, boron and aluminum. Aluminum forms the oxide Al_2O_3, which reacts with water indirectly to form the compound $Al(OH)_3$, which acts as a base toward strong acids and as an acid toward strong bases.

Boron is a non-metal and forms the oxide B_2O_3, which combined with water forms the compound $B(OH)_3$. This com-

pound has only acid properties and therefore the formula is usually written H_3BO_3 and is known as boric acid.

The Fourth Group.

Carbon forms two oxides. The first, CO, forms neither acids nor bases. The second, CO_2, unites with water to form an unstable acid, $CO(OH)_2$, called carbonic acid and commonly written H_2CO_3.

Silicon forms one oxide, SiO_2, which forms a number of acids when combined with varying amounts of water.

$SiO_2 + H_2O = SiO(OH)_2$ or H_2SiO_3, called meta-silicic acid.
$SiO_2 + 2H_2O = Si(OH)_4$ or H_4SiO_4, called ortho-silicic acid.

Tin forms two oxides, SnO, basic, and SnO_2, both basic and acidic.

$SnO_2 + H_2O = SnO(OH)_2$ or H_2SnO_3, stannic acid.

Lead forms several oxides, only one of which, PbO, need be considered here. This oxide unites with water to form $Pb(OH)_2$, which is basic toward acids and weakly acidic toward strong bases.

The Fifth Group.

Nitrogen forms several oxides, three of which are acid producers.

$N_2O + H_2O = 2HNO$, Hyponitrous acid.
$N_2O_3 + H_2O = 2HNO_2$, Nitrous acid.
$N_2O_5 + H_2O = 2HNO_3$, Nitric acid.

Phosphorus forms two oxides each of which forms acids with water.

$P_2O_3 + 3H_2O = 2P(OH)_3$ or $2H_3PO_3$, phosphorus acid.
$P_2O_5 + H_2O = 2HPO_3$, meta-phosphoric acid.
$P_2O_5 + 2H_2O = H_4P_2O_7$, pyro-phosphoric acid.
$P_2O_5 + 3H_2O = 2H_3PO_4$, ortho-phosphoric acid.

Arsenic and antimony form oxides and acids similar to those of phosphorus.

Bismuth forms several oxides, only one of which need be considered here.

Bi_2O_3 is a basic oxide forming bismuth hydroxide with water.

The Sixth Group.

Sulphur forms two oxides, SO_2 and SO_3, both acidic.

$SO_2 + H_2O = H_2SO_3$, sulphurous acid.
$SO_3 + H_2O = H_2SO_4$, sulphuric acid.

Chromium does not form the oxide CrO but the corresponding hydroxide, $Cr(OH)_2$, is well known. Cr_2O_3, chromic oxide, is weakly basic and forms chromic hydroxide, $Cr(OH)_3$.

CrO_3, chromium trioxide or chromic anhydride, is acidic and forms chromic acid, $CrO(OH)_2$ or H_2CrO_4.

The Seventh Group.

This group with the exception of manganese may best be considered as forming a regular series of acids corresponding to the four regular valencies. The exceptions will be learned in the study of the separate elements.

$Cl_2O + H_2O = 2HClO$, hypochlorous acid.
$Cl_2O_3 + H_2O = 2HClO_2$, chlorous acid.
$Cl_2O_5 + H_2O = 2HClO_3$, chloric acid.
$Cl_2O_7 + H_2O = 2HClO_4$, per-chloric acid.

Bromine does not form any oxides but the acids HBrO, hypobromous, and $HBrO_3$, bromic, are known. Iodine may be considered as forming acids similar to the chlorine acids. Only one oxide has been prepared in the free condition, I_2O_5.

Manganese forms six oxides, some of which are basic and some acidic.

MnO, manganous oxide, basic, forming $Mn(OH)_2$.
Mn_2O_3, manganic oxide, forming $Mn(OH)_3$, basic.
MnO_3, manganese trioxide or manganic anhydride.
$MnO_3 + H_2O = H_2MnO_4$, manganic acid.
Mn_2O_7, manganese hept-oxide or per-manganic anhydride.
$Mn_2O_7 + H_2O = 2HMnO_4$, per-manganic acid.

The Eighth Group.

The important oxides of this group correspond to the valencies of two and three. They are all basic and form basic hydroxides.

70. The System of naming Acids and Salts.

The names of salts formed from "ous" acids end in "ite," and the names of salts formed from "ic" acids end in "ate."

If an acid has less oxygen than the usual "ous" acid, that is if it represents a lower stage of oxidation, the prefix "hypo" is placed before the name of the "ous" acid. Thus HNO_2 is nitrous acid, HNO is hyponitrous acid.

If an acid represents a higher stage of oxidation than the usual "ic" acid, the prefix "per" is placed before the name. Thus, $HClO_3$ is chloric acid and $HClO_4$ is per-chloric acid.

The Names and Valencies of Radicals.

The names of radicals are taken as the first or last part of the name of a salt. Thus,

$NaNO$ is sodium hypo-nitrite, NO is the hypo-nitrite radical and has a valence of one because it can form a compound with one hydrogen.

$Mg_2P_2O_7$ is magnesium pyro-phosphate, P_2O_7 is the pyro-phosphate radical and has a valence of four because it can exist in combination with four hydrogen atoms.

OXIDATION IN THE WET WAY.

71. Many oxidations can best be carried out in a solution containing the required oxidizing agent and the substance to be oxidized.

In such reactions no oxygen is actually set free in the gaseous form but enters at once into combination with the reducing agent. If there is an excess of the reducing agent all the oxidizing agent will be used up. If there is an excess of the oxidizer, all the reducing agent will be transformed. If any basic oxides are formed in the reaction an acid must be present to change them to the corresponding salts.

Chlorine, bromine and iodine oxidize in the presence of water by virtue of extracting the hydrogen of the water, forming the corresponding acid, and liberating oxygen.

$$2Cl_2 + 2H_2O = 4HCl + O_2.$$
$$2Br_2 + 2H_2O = 4HBr + O_2.$$
$$2I_2 + 2H_2O = 4HI + O_2.$$

Compounds containing oxygen and two other elements may be considered as made up of two oxides. If the compound is an acid, one of the oxides will be water and the other an acid anhydride. If it is a salt, one of them will be a basic oxide and the other an acid anhydride.

In either case, if the acid anhydride is an oxide of an element corresponding to a high combining power, and if an oxide corresponding to a lower combining power exists, the substance is an oxidizing agent.

Consider sulphuric acid, H_2SO_4.

$H_2SO_4 = H_2O + SO_3$, but SO_3 can give up oxygen and drop to SO_2. Therefore we have

$$H_2SO_4 = H_2O + SO_2 + O.$$

$$2K_2CrO_4 = 2K_2O + 2CrO_3.$$

CrO_3 can give up oxygen and become Cr_2O_3; so we have

$$2K_2CrO_4 = 2K_2O + Cr_2O_3 + 3O.$$

$$2KMnO_4 = K_2O + Mn_2O_7.$$

Mn_2O_7 can drop to $2MnO$; therefore we have,

$$2KMnO_4 = K_2O + 2MnO + 5O.$$

Exercise.

72. State the formulas of ten different oxidizing agents and indicate, by equations, how they oxidize.

Formation of Equations for Oxidation Reactions in the Wet Way.

73. It is required to oxidize phosphorus with nitric acid. First, how does nitric acid oxidize?

$2HNO_3 = H_2O + 2NO + 3O$, that is, when nitric acid is used as an oxidizing agent, two molecules give three atoms of oxygen, and nitric oxide is given off.

When phosphorus is oxidized, two atoms require five atoms of oxygen to form the pentoxide. Oxygen is furnished by the nitric acid in multiples of three and needed by the phosphorus in multiples of five. Therefore we shall have to use five times

two molecules of nitric acid and three times two atoms of phosphorus.

$$10HNO_3 = 5H_2O + 10NO + 15O.$$
$$6P + 15O = 3P_2O_5.$$

Phosphorus pentoxide is an acid anhydride and will combine with the proper quantity of water to form phosphoric acid.

$$3P_2O_5 + 9H_2O = 6H_3PO_4.$$

Adding these equations together, we have the final equation for this reaction.

$$4H_2O + 6P + 10HNO_3 = 6H_3PO_4 + 10NO.$$

Exercise.

74. Indicate the oxidation of the other elements of the fifth group, in the same way.

75. If the oxidizing agent is a substance that gives rise to a basic oxide in decomposing, there must be an acid present to combine with this oxide and form the corresponding salt.

Consider manganese dioxide with hydrochloric acid.

$$MnO_2 = MnO + O. \quad MnO \text{ is a basic oxide.}$$

One atom of O will combine with two atoms of H.

$$O + 2HCl = H_2O + Cl_2.$$

The MnO requires two molecules of hydrochloric acid.

$$MnO + 2HCl = MnCl_2 + H_2O.$$

Adding these equations together, we have

$$MnO_2 + 4HCl = MnCl_2 + 2H_2O + Cl_2.$$

Exercise.

76. Indicate the oxidation of the other halogen acids in the same way. Also the oxidation of H_2SO_3, sulphurous acid, to H_2SO_4.

Suppose we had to oxidize hydrochloric acid with potassium permanganate, $KMnO_4$.

$$2KMnO_4 = K_2O + 2MnO + 5O.$$

5O take 10 H

$$5O + 10HCl = 5H_2O + 5Cl_2.$$

K_2O takes 2 HCl

$$K_2O + 2HCl = 2KCl + H_2O.$$

2MnO takes 4 HCl

$$2MnO + 4HCl = 2MnCl_2 + 2H_2O.$$

Adding these equations together, we have the final equation:

$$2KMnO_4 + 16HCl = 2MnCl_2 + 2KCl + 8H_2O + 5Cl_2.$$

EXAMPLES OF OXIDATION INVOLVING A CHANGE FROM A LOWER TO A HIGHER STAGE OF OXIDATION.

77. Change cuprous chloride, CuCl, to cupric chloride, $CuCl_2$. Add chlorine directly,

$$CuCl + Cl = CuCl_2.$$

Or add hydrochloric acid and an oxidizing agent. Use nitric acid. Two molecules of this acid give three atoms of oxygen for oxidizing purposes; three atoms of oxygen are equivalent to six atoms of hydrogen or to six molecules of hydrochloric acid. So the least quantity of chlorine that we can use will be six atoms. We must therefore oxidize six molecules of cuprous chloride.

Writing all the equations and adding them together, we have

$$\begin{aligned} 2HNO_3 &= H_2O + 2NO + 3O. \\ 3O + 6HCl &= 3H_2O + 3Cl_2. \\ 3Cl_2 + 6CuCl &= 6CuCl_2. \end{aligned}$$

$$6CuCl + 2HNO_3 + 6HCl = 6CuCl_2 + 2NO + 4H_2O.$$

Indicate this same oxidation, using different oxidizing agents.

Suppose it is required to oxidize ferrous sulphate to ferric sulphate by means of potassium permanganate.

The formula of ferrous sulphate is $FeSO_4$, and for ferric sulphate is $Fe_2(SO_4)_3$. It is evident that two molecules of the ferrous salt will be required to make one of the ferric.

The basic oxide contained in the ferrous sulphate is FeO, and in the ferric sulphate is Fe_2O_3. For each two molecules of the first compound therefore there must be one atom of oxygen added, as well as some more SO_4.

Potassium permanganate oxidizes in acid solution according to the equation,

$$2KMnO_4 = K_2O + 2MnO + 5O.$$

Since five oxygen atoms are available, at least ten molecules of ferrous sulphate must be used, in which case five molecules of sulphuric acid will be necessary to furnish the five additional SO_4 radicals.

$$10FeSO_4 + 2KMnO_4 + 5H_2SO_4 = 5Fe_2(SO_4)_3 + K_2O + 2MnO + 5H_2O.$$

$$2MnO + 2H_2SO_4 = 2MnSO_4 + 2H_2O.$$

$$K_2O + H_2SO_4 = K_2SO_4 + H_2O.$$

$$10FeSO_4 + 2KMnO_4 + 8H_2SO_4 = 5Fe_2(SO_4)_3 + 2MnSO_4 + K_2SO_4 + 8H_2O.$$

SOME OXIDIZING AGENTS WITH EQUATIONS SHOWING HOW THEY OXIDIZE IN THE WET WAY.

78. 1. $PbO_2 = PbO + O$.

2. $MnO_2 = MnO + O$, in the presence of a reducing agent and an acid.

3. $Cl_2 + H_2O = 2HCl + O$, in the presence of a reducing agent.

4. $Br_2 + H_2O = 2HBr + O$, in the presence of a reducing agent.

5. $I_2 + H_2O = 2HI + O$, in the presence of a reducing agent.

6. $KClO_3 = KCl + 3O$, when heated to a high temperature.

7. $Pb_3O_4 = 3PbO + O$, when heated or in the presence of an acid and a reducing agent.

8. $2HNO_3 = H_2O + 2NO + 3O$, in the presence of a reducing agent.

9. K_2CrO_4, potassium chromate, $2K_2CrO_4 = 2K_2O + Cr_2O_3 + 3O$, in the presence of an acid and a reducing agent.

10. $K_2Cr_2O_7$, potassium bichromate, dichromate or acid chromate. $K_2Cr_2O_7 = K_2O + Cr_2O_3 + 3O$, in the presence of an acid and a reducing agent.

11. $KMnO_4$, potassium permanganate, $2KMnO_4 = K_2O + 2MnO + 5O$, in the presence of a reducing agent and an acid.

12. K_2MnO_4, potassium manganate. $K_2MnO_4 = K_2O + MnO + 2O$, in the presence of an acid and a reducing agent.

Exercise.

79. Indicate, by equations, the following oxidations, making use in each case of the above oxidizing agents.

C, carbon, to CO_2, carbon dioxide. **68.**
CO, carbon monoxide, to CO_2, carbon dioxide.
H, hydrogen, to H_2O, water or hydrogen oxide.
CH_4, methane, to $CO_2 + H_2O$.
C_2H_5OH, alcohol, to $CO_2 + H_2O$.
C_2H_5OH to CH_3CHO, aldehyde.
C_2H_5OH to CH_3COOH, acetic acid.
S, sulphur, to SO_2, sulphur dioxide.
S, sulphur, to SO_3, sulphur trioxide.
H_2S, hydrogen sulphide, to $S + H_2O$.
H_2S, hydrogen sulphide, to H_2SO_4.
HCl, hydrogen chloride, to $H_2O + Cl$. **75.**
HBr, hydrogen bromide, to $H_2O + Br$. **75.**
HI, hydrogen iodide, to $H_2O + I$. **75.**
HI, hydrogen iodide, to HIO_3, iodic acid.
$H_2C_2O_4$, oxalic acid, to $CO_2 + H_2O$.
$C_{17}H_{35}CO_2H$, stearic acid, to $CO_2 + H_2O$.
$FeCl_2$, ferrous chloride, to $FeCl_3$, ferric chloride. **77.**
$SnCl_2$, stannous chloride, to $SnCl_4$, stannic chloride. **77.**
$HgCl$, mercurous chloride, to $HgCl_2$, mercuric chloride. **77.**
$FeSO_4$, ferrous sulphate, to $Fe_2(SO_4)_3$, ferric sulphate. **77.**
P, phosphorus, to H_3PO_4, phosphoric acid. **73.**
P_2O_3, phosphorus trioxide, to H_3PO_4, phosphoric acid.
As, arsenic, to H_3AsO_4, arsenic acid. **73.**
As_2O_3, arsenic trioxide, to H_3AsO_4, arsenic acid.

REDUCTION REACTIONS.

80. **Reduction** is exactly the opposite of oxidation. The oxidizing agent oxidizes the reducing agent, and the reducing agent reduces the oxidizing agent.

Any substance that will combine directly with oxygen is necessarily a **Reducing agent**; or any substance that will furnish or give up anything that will combine with oxygen.

Some common reducing agents:

H, C, CO, Na, Mg, Al, Fe, H_2S, HI, SO_2, H_2SO_3, sulphurous acid, $SnCl_2$, stannous chloride, CH_4, methane, $H_2C_2O_4$, oxalic acid, $FeCl_2$, ferrous chloride, Zn, $FeSO_4$, ferrous sulphate.

81. Examples of Reduction Reactions.

$$Fe_2O_3 + 3H_2 = 2Fe + 3H_2O.$$
$$2FeCl_3 + Fe = 3FeCl_2.$$
$$2FeCl_3 + H_2S = 2FeCl_2 + 2HCl + S.$$
$$2H_2SO_4 + C = 2H_2O + 2SO_2 + CO_2.$$
$$4Cu_2O + CH_4 = 8Cu + CO_2 + 2H_2O.$$

Exercise.

82. Show the reduction of the various metallic oxides, each with H, C and CO.

Reduce $CuCl_2$ to CuCl with copper.

Reduce $HgCl_2$ to HgCl with mercury.

Reduce $HgCl_2$ to HgCl with stannous chloride.

Reduce $FeCl_3$ to $FeCl_2$ with H_2S, $SnCl_2$ and H_2SO_3.

Reduce $KMnO_4$ to K_2SO_4 and $MnSO_4$ with oxalic acid, in the presence of sulphuric acid.

Reduce $K_2Cr_2O_7$ to K_2SO_4 and $Cr_2(SO_4)_3$, in the presence of sulphuric acid, with stearic acid, $C_{17}H_{35}COOH$.

CHEMICAL EQUIVALENCE BETWEEN COMPOUNDS.

83. One molecule of hydrogen phosphate, H_3PO_4, is equivalent to three molecules of hydrochloric acid, nitric acid, sodium hydroxide or sodium chloride.

Two molecules of hydrogen phosphate are equivalent to three molecules of sulphuric acid, calcium chloride or barium hydroxide.

One molecule of sulphuric acid is equivalent to two molecules of hydrochloric acid, nitric acid, sodium chloride or sodium hydroxide.

One molecule of sulphuric acid is equivalent to one molecule of calcium nitrate, zinc sulphate or sodium carbonate.

Three molecules of nitric acid are equivalent to one molecule of hydrogen phosphate; two molecules to one molecule of sul-

phuric acid, and one molecule to one molecule of sodium hydroxide.

A certain quantity of one chemical compound is said to be equivalent to a definite quantity of another compound, when they may both be completely involved in the same reaction or can each react with a fixed quantity of another substance.

In general, the quantities represented in any complete equation are chemically equivalent for that particular reaction.

PART III.

MORE ADVANCED THEORY.

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84. **Avogadro's Hypothesis.** *Equal volumes of all gases at the same temperature and pressure contain the same number of molecules.*

Or in other words, all gaseous molecules under the same conditions occupy equal space.

If an equation has been properly written so that each symbol represents a molecule, then from the above principle, the relative volumes of the various substances entering that reaction, when considered to be in the gaseous form, are proportional to the number of molecules as indicated by the coefficients in the equation.

Example. How many liters of oxygen would be required to burn ten liters of alcohol vapor?

$$C_2H_5OH + 3O_2 = 2CO_2 + 3H_2O.$$

This equation means that one molecule of alcohol requires three molecules of oxygen to burn it and results in the formation of two molecules of carbon dioxide and three molecules of water.

In other words, ten liters of alcohol vapor require thirty liters of oxygen to burn it and form twenty liters of carbon dioxide and thirty liters of water vapor.

Exercise.

85. Write equations showing the oxidation or burning of gases and show how many liters of the various products would be formed if ten liters of the gas were used. 68.

MOLECULAR AND ATOMIC WEIGHTS.

86. **The hydrogen molecule is divisible.** If one volume of hydrogen be combined with one volume of chlorine, two volumes of hydrochloric acid gas are formed. By Avogadro's Hypothesis there must be twice as many molecules of the compound as there are of either of the elementary gases. Therefore the hydrogen molecule must split up into at least two particles. The same is true of the chlorine molecule.

87. **The molecule of hydrochloric acid gas must contain equal numbers of hydrogen and chlorine atoms.** When the compound is analyzed it is found to contain one part of hydrogen to thirty-five and one-half parts of chlorine by weight. Since the atomic weight of hydrogen is one and that of chlorine thirty-five and one-half, the formula must be H_nCl_n.

88. **The molecule of hydrogen chloride contains but one hydrogen atom.** If it contained two or more hydrogen atoms a part of this hydrogen could be replaced by a metal, forming an acid salt having the general formula $NaH_{n-1}Cl_n$. Lab. Ex. No. 25. Or if it contained two hydrogen atoms a compound having the formula $HNaCl_2$ would have been found. No such compound has ever been made.

89. **The molecule of hydrogen contains two atoms.** Since it has been demonstrated that the formula of hydrochloric acid gas is HCl, and since one volume or one molecule of hydrogen gives two volumes or two molecules of hydrochloric acid gas, each molecule of hydrogen must contain only two atoms.

90. **Gas Density** is the ratio of the weight of a certain volume of a gas to the weight of an equal volume of hydrogen at the same temperature and pressure. Sometimes the ratio of the weight of a certain volume of the gas to the weight of an equal volume of oxygen is taken and the result multiplied by sixteen.

91. **Vapor density** is an expression that conveys the same idea as gas density but is more properly used in reference to substances that do not ordinarily exist in the gaseous form.

92. **The molecular weight of a substance is equal to twice its gas density.** The ratio of the weight of a certain volume of a gas to the weight of an equal volume of hydrogen is the same

as the ratio of the weight of a molecule of the gas to the weight of a hydrogen molecule, since equal volumes of all gases under the same conditions contain the same number of molecules. In other words, gas density is a number that tells how many times heavier a molecule is than one molecule of hydrogen. Since the molecule of hydrogen contains two atoms and therefore has a molecular weight of two, twice the gas density must be a number that tells how many times heavier a molecule of a gas is than one atom of hydrogen. But this is the definition of molecular weight.

93. **The determination of gas density.** A glass balloon similar to that indicated in fig. 17 and having a capacity vary-

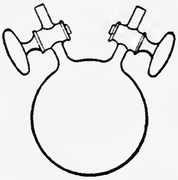

Fig. 17.

ing from 100 c.c. to 1000 c.c. is filled with a gas and weighed. It is then filled with hydrogen and weighed again. These weights are evidently not the true weights of the gases, but the weight of the gas plus the weight of the bulb minus the buoyant effect of the air in which the weighing was made. By means of the mercury pump all the gas is now removed from the bulb, and it is weighed again. This weight is the weight of the bulb less the buoyant effect of the air. If this weight be subtracted from the first two weights, the real weights of the two gases will be obtained. By dividing the weight of the gas by the weight of the hydrogen the gas density will be found in terms of hydrogen. Since the ratio of the atomic weight of oxygen to hydrogen is not exactly 16 to 1, and since tables based on oxygen as 16 are in common use, it is better to compare the weight of the unknown gas to the weight of an equal volume of oxygen and multiply the result by 16. This gives the gas density in terms of one-sixteenth of the oxygen atom.

94. **Determination of vapor density. Meyer's method.** In this method a small quantity of the substance is vaporized, and the volume of the air displaced by the vapor is measured.

This volume is reduced to the corresponding volume at 0° C. temperature and 760 mm. pressure and multiplied by the weight of 1 c.c. of oxygen under the same conditions. **105.** The ratio of the weight of the substance taken, to this weight, shows how many times heavier the molecule is than one molecule of oxygen. This result multiplied by sixteen gives the vapor density of the material. The apparatus used is indicated in fig. 18. It consists of two glass tubes. The outer one has a bulb at the bottom in which is placed some liquid that has

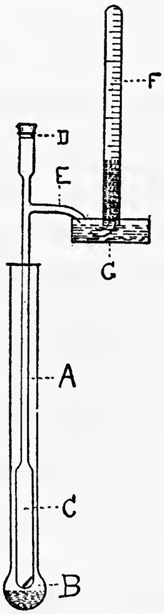

Fig. 18.

a higher boiling point than that of the substance to be vaporized. The vapor density tube is suspended in this and the liquid in the outer tube heated to boiling. After the air in the inside tube has ceased to expand, the end of the delivery tube at E is placed under an inverted measuring tube, F, full of water. About a tenth of a gram of the substance is placed in a small glass stoppered vial and carefully weighed. The cork at D is removed, the vial dropped in and the cork quickly replaced. The vial passes through the long inner tube to the enlarged

portion at C. The heat causes the material to vaporize, the stopper is blown out, and a volume of air escapes into the graduated tube, F, just equal to the volume of the vapor.

95. **Determination of the formula from the analysis and vapor density.** Suppose a new substance is found and it is desired to determine its formula. It is first analyzed and its chemical composition ascertained in per cents. Then its vapor density is found by the above methods or by other methods to be described later. **110, 111** and **112.**

If each percentage be divided by the corresponding atomic weight a series of numbers will be obtained which will be proportional to the numbers of the different atoms in the molecule.

If this series be divided by the smallest number, the results expressed as fractions and then reduced to a common denominator, the numerators will be the numbers of atoms of each element corresponding to the simplest formula.

Example. Suppose a substance contained carbon, 33.70%; hydrogen, 3.38%; oxygen, 62.92%; what would be the formula corresponding?

Dividing these per cents. by the corresponding atomic weights we have

$$\frac{33.70}{12}=2.81; \quad \frac{3.38}{1}=3.38; \quad \frac{62.92}{16}=3.93.$$

This means that the formula must be

$$C_{2.81}\ H_{3.38}\ O_{3.93}.$$

Dividing each of these sub-numbers by the smallest and expressing the results as fractions, we have

$$C_1\ H_{1\frac{1}{5}}\ O_{1\frac{2}{5}}.$$

This means that there are one and one-fifth as many hydrogen atoms, and one and two-fifths as many oxygen atoms as there were carbon atoms. This could only be the case if the formula were

$$C_5H_6O_7.$$

If the molecular weight corresponding to this formula agrees with that found from the vapor density, then this formula is the correct one. Otherwise it is the simplest possible formula, and some multiple of it will correspond to the molecular weight called for by the vapor density.

If the vapor density determination is exact, the molecular weight determined from it may be multiplied at once by the various percentages, and the results divided by the atomic weights will give the correct formula.

Exercise.

96. Determine the simplest formulas corresponding to the following analyses:

1. Lead, 86.60%, Sulphur, 13.40%.
2. Calcium, 51.10%, Fluorine, 48.90%.
3. Mercury, 84.90%, Chlorine, 15.10%.
4. Arsenic, 75.80%, Oxygen, 24.20%.
5. Hydrogen, 1.59%, Nitrogen, 22.22%, Oxygen, 76.19%.
6. Iron, 70.00%, Oxygen, 30.00%.

Exercise.

97. Determine the correct formulas corresponding to the following analyses and vapor densities:

	Carbon.	Hydrogen.	Nitrogen.	Oxygen.	Vapor Density.
1.	91.30%	8.70%	....	....	45.90
2.	85.71%	14.29%	....	....	35.00
3.	54.55%	9.09%	....	36.36%	66.00
4.	40.00%	6.67%	....	53.33	30.00
5.	92.30%	7.70%	....	....	26.00
6.	77.32%	7.63%	15.05%	....	46.50

DETERMINATION OF ATOMIC WEIGHTS.

98. **An atom is the smallest particle of an element that is known to form part of a molecule.**

An **atomic weight** is the ratio of the weight of the smallest particle of an element forming part of a molecule to the weight of one atom of hydrogen.

To determine the atomic weight of an element it is first necessary to analyze a series of its compounds and determine the percentage of the element present in each.

The percentage of the element present in a single molecule will of course be the same as that found in the quantity of material analyzed. Therefore if the molecular weight of each substance be multiplied by the percentage of the element found

in it, the result will be the total weight of that element in the molecule. The smallest of these weights will be the required atomic weight.

For example it is desired to find the atomic weight of chlorine. A series of compounds of chlorine is analyzed, and the percentage of that element present in each is determined. Then the molecular weights of all the substances are determined.

The molecular weight of each is multiplied by the percentage of chlorine. The result is the number of weights of chlorine in each molecule. An inspection of the results shows at once what the atomic weight of the element is.

The following table illustrates this.

Substance.	Vapor density.	Mol. weight.	Per cent. of chlorine.	Weight present in molecule.
Hydrochloric acid........	18.25	36.50	97.26%	35.5
Methyl chloride..........	25.25	50.50	70.30	35.5
Methylene chloride.......	49.50	99.00	71.72	71.0
Mercuric chloride........	135.50	271.00	26.20	71.0
Chloroform..............	59.25	118.50	89.87	106.5
Carbon tetrachloride.....	77.00	154.00	92.21	142.0

Since the methods of obtaining molecular weights are not absolutely exact, the atomic weights found as above are not exact. The correct weights are found by the analysis of compounds that admit of being prepared in a condition of great purity. See also **127.**

THE NATURE OF GASEOUS BODIES.

99. All gaseous bodies, that is, gases and vapors, are supposed to be made up of an almost infinite number of molecules traveling in straight lines with enormous velocity. Their directions are constantly altered by interference with each other or with the walls of the inclosing vessel, but they are always deflected in straight lines. The path of a single molecule is therefore made up of a continuous broken line, the straight portions of which are very long compared with the size of the

molecules. **The velocity of gaseous molecules of a given kind depends** on the temperature alone. If the temperature be raised, the molecular velocity is increased. The temperature of a gas is proportional to the average molecular velocity.

The relative velocity of gaseous molecules can be determined by measuring the relative rates at which different gases escape through very small openings in very thin walls. From this it appears that the velocities are to each other **inversely as the square roots of the gas densities.**

Gaseous pressure is due to the multitude of impacts of the molecules against the walls of the inclosing vessel.

100. **Boyle's Law.** *At constant temperature the volume of a gas varies inversely as the pressure.* That is, if the volume be reduced to one-half of what it was, the pressure will be doubled.

This corresponds exactly with the above theory, for if the volume be decreased to one half of what it was, there will be twice as many molecules per unit of volume and consequently twice as many striking the walls and producing pressure as there were before.

This law may be more generally expressed by the equation,

$$\frac{P}{P'} = \frac{V'}{V} \text{ or } PV = P'V'$$

in which P and P′ are the two different pressures to which a body of gas is subjected at constant temperature, and V and V′ are the corresponding volumes.

101. **Effect of temperature.** If a volume of gas at 0° C. be heated sufficiently to increase its temperature one degree, its volume will be increased $\frac{1}{273}$ of what it was at 0° C. And for each additional change of one degree the volume will change $\frac{1}{273}$ of what it was at 0° C. Therefore if the temperature of a volume of gas at 0° C. be decreased 273 degrees, the volume would theoretically become zero.

This temperature, 273 degrees below zero, is known as **absolute zero.** Temperatures reckoned from this point are known as **absolute temperatures.** Common temperatures are changed to absolute temperatures by adding 273.

102. **The Law of Charles.** *At a constant pressure the volume of a body of gas varies directly as its absolute temperature.*

This law may be more generally expressed by the equation, $\frac{V}{V'}=\frac{T}{T'}$ or $VT'=V'T$, in which V and V′ are the volumes of a body of gas at constant pressure, corresponding to the absolute temperatures T and T′.

Given the Volume of a Body of Gas at known Temperature and Pressure, to compute the Corresponding Volume, Temperature or Pressure under other known Conditions.

103. Let V be the given volume at the temperature T and pressure P. Assuming that the temperature remains constant, let V′ be the new volume corresponding to the new pressure P°. Then we have,

$$PV=P^{\circ}V' \text{ or } V'=\frac{PV}{P^{\circ}}.$$

Now let us assume that the pressure remains constant and that the temperature be allowed to become T°. The gas will assume a new volume, which we will call V°. Applying the law of Charles we have, $V'T^{\circ}=V^{\circ}T$; substituting the value of V′ from above we have,

$$\frac{PV}{P^{\circ}}T^{\circ}=V^{\circ}T, \text{ or } \frac{PV}{T}=\frac{P^{\circ}V^{\circ}}{T^{\circ}}.$$

This is the general gas law. P, V and T will always be known, also two of the remaining factors, making it a very simple matter to determine the remaining one.

104. **Example.** The volume of a body of gas is measured at a temperature of 22° and a pressure of 780 mm. and found to be 25 c.c., what would the corresponding volume be at a temperature of 100° and a pressure 1,000 mm.? Substituting in the above formula we have,

$$\frac{25\times 780}{295}=V\frac{1000}{373} \text{ or } V=\frac{25\times 780\times 373}{295\times 1000}.$$

Note. Since the volumes of gases are always measured subject to atmospheric pressure, it is customary to speak of the pressures of millimeters of mercury; the actual pressure being

that due to a column of mercury a certain number of millimeters high.

105. **Normal Temperature and Pressure.** Measured volumes of gases are generally reduced to the corresponding volume at 0° and 760 mm. pressure, known as **Normal Temperature and Pressure.**

$$\frac{PV}{T} = \frac{P^\circ V^\circ}{T^\circ} = \frac{760}{273} V^\circ$$

$$V^\circ = \frac{273}{760} \times \frac{P}{t+273} V = .35921 \frac{PV}{273+t}$$

in which V° is the required volume and t the observed temperature, and P the actual pressure to which the gas is subjected at the time the volume V is read: It is equal to the barometric pressure minus the vapor pressure of the liquid in contact with the gas, at the temperature t°.

Exercise.

106. 1. Change any given volume of a gas from any given temperature and pressure to the corresponding volume at any other temperature and pressure.

2. Compute the temperature or the pressure of a volume of a body of gas at known pressure and temperature corresponding to a required volume and temperature or volume and pressure.

3. Reduce a given volume of gas from known temperature and pressure to normal temperature and pressure.

EVAPORATION AND CONDENSATION.

107. **Evaporation and the Nature of a Vapor.** At a constant temperature, a liquid in an inclosed space evaporates until the number of molecules escaping per unit of time is equaled by the number of gaseous molecules reëntering the liquid condition. A definite condition of equilibrium is then established.

The rate of evaporation increases with increase of temperature and decrease of pressure. The presence of substances dissolved in the liquid not only decreases the tendency for vapor to form, but may cause the vapor to condense to the liquid form.

A vapor that is in equilibrium with its liquid is said to be saturated. If the pressure due to a saturated vapor in contact with its liquid be suddenly reduced, vapor will form in the liquid at other points than on the surface, and the liquid is said to boil.

Water evaporating at the sea level in contact with the atmosphere is constantly subject to a pressure of about fifteen pounds to the square inch. If the temperature be increased the vapor pressure will finally become great enough to lift the atmosphere away from the surface and the liquid will boil. The **boiling point** is the temperature at which the vapor pressure equals the atmospheric pressure. Water boils at low temperatures on high mountains, and will boil at ordinary temperatures in a vacuum.

A saturated vapor at a given temperature has a definite number of molecules per unit volume. These travel in straight lines like the molecules of a gas and exert pressure on the walls of the inclosing vessel. If the volume be decreased at constant temperature, some of the vapor enters the liquid condition, and the number of molecules per unit volume remains unchanged. In other words, the volume of a saturated vapor may be decreased without an increase of pressure.

108. **Condensation and Liquefaction.** A liquid in an inclosed space at a given temperature is always in contact with its saturated vapor. This vapor exerts a definite pressure which is dependent on the temperature alone.

If the pressure be kept constant and the temperature reduced, the vapor will constantly change to the liquid form, or condense. If the temperature be maintained constant and the pressure increased, condensation will again take place.

109. **Critical Temperature and Pressure.** If a liquid be contained in an inclosed space in contact with its saturated vapor and the temperature be increased, the liquid will evaporate and the pressure of the vapor will increase. This will continue until the density of the vapor is the same as that of the liquid; and the surface of the liquid, that is, the line of demarkation between the vapor and liquid, will disappear, the two becoming identical. The temperature at which this takes place is known as the **Critical Temperature.** In other words, the critical temperature is that above which no amount of compression will cause the gas to condense. **Critical Pressure** is

the pressure of a vapor just sufficient to condense it at its critical temperature.

Effect of Substances in solution on the Boiling and Freezing Point.

110. The vapor pressure of a solution of a substance is less than that of the pure solvent. This is because the presence of the molecules of the dissolved substance not only tends to prevent the escape of molecules of the solvent but actually attracts some of those that have escaped, back into the liquid.

If a solution of some substance in water be placed in the same inclosed space with some water, and the temperature be constant, the water will entirely evaporate and condense in the solution. This is because the vapor pressure over the solution can never become as great as the pressure over the water. Therefore equilibrium cannot be established until the water has entirely evaporated. As the space was inclosed, this can only take place as a result of the condensation of the vapor into the solution.

111. **The Boiling Point Law.** If different portions of the same solvent contain the same number of particles of dissolved substances per unit volume, the boiling point will be raised by the same amount.

Assuming for the present that there is no dissociation in solution, this gives a very simple method of getting molecular weights. It is only necessary to find the elevation of the boiling point produced by a gram molecule of a known substance in a suitable solvent, then add to an equal volume of the same solvent sufficient of the unknown substance to produce the same elevation of the boiling point. When this has been done one gram molecule of the unknown substance must have been used.

112. **The Freezing Point Law.** Equal numbers of particles per unit volume in different portions of the same solvent produce equal depressions of the freezing point. This again gives a method for determining molecular weights. It is only necessary to compare the depression of the freezing point produced by a substance whose molecular weight is known with that produced by a definite weight of a substance whose molecular weight is desired.

Exercise.

113. What will be the final condition of equilibrium in each of the following cases? Give the full explanation.

1. A liquid in an inclosed space at a constant temperature.
2. Two portions of the same liquid in an inclosed space at a constant temperature, one portion being above the other.
3. Two portions of the same liquid contained in the same inclosed space, at different constant temperatures.
4. A water solution of some substance in the same inclosed space with a vessel of water at constant temperature.

OSMOTIC PRESSURE.

114. Dissolved substances exert a pressure in solution which is exactly analogous to gaseous pressure. This pressure is, however, not appreciable, unless the solution is contained in a vessel provided with a semi-permeable membrane, and the whole immersed in water. Such a membrane allows the passage of the water, but will not allow the passage of molecules of dissolved substances.

Suppose some sugar solution were placed in a rather large glass tube, the bottom of which had been closed with a semi-permeable membrane and the lower end placed in a jar of water. The solution would rise in the tube, and water would pass from the jar through the membrane into the tube until the pressure due to the column of liquid was equal to the osmotic pressure of the diluted solution.

Again, suppose one gram-molecule of sugar, 342 g., be placed in a semi-permeable cell of one liter capacity, the whole immersed in water and the osmotic pressure measured. It would be found to be 22.4 atmospheres. At a higher temperature the pressure would become higher. If the sugar were considered as a gas, one liter of the gas would weigh ½ its molecular weight times .0896, which would be 15.32 g. 342 grams of sugar in the gaseous form at normal temperature and pressure would therefore occupy 22.4 liters. If this volume were condensed to one liter the pressure would be 22.4 atmospheres.

The Laws of Osmotic Pressure.

115. It should be noted that these laws are identical with the laws of Boyle, Charles and Avogadro.

1. With a fixed quantity of material in solution, at a constant temperature, the osmotic pressure varies inversely as the volume.

2. With a fixed volume and quantity of material, the osmotic pressure is proportional to the absolute temperature.

3. Equal volumes of the same liquid at the same temperature and showing equal osmotic pressures contain equal numbers of particles.

Computation of Osmotic Pressures.

116. One gram-molecule of any substance in the gaseous form at normal temperature and pressure occupies 22.4 liters.

The problem is to find what the pressure would be if one gram-molecule or fraction thereof of any material is placed in an osmotic cell at a given temperature.

The general gas law may be utilized.

$$\frac{P^\circ V^\circ}{T^\circ} = \frac{PV}{T}.$$

P° is a pressure of one atmosphere.

V° is the volume of the gram-molecule at normal temperature and pressure equal to 22.4.

T° is 273° absolute.

The expression then becomes,

$$\frac{PV}{T} = \frac{22.4}{273}.$$

P is the osmotic pressure, V the volume of the cell and T the absolute temperature.

$$P = \frac{T}{V} \times \frac{22.4}{273}.$$

Example. What would be the osmotic pressure of 114 grams of sugar in one liter of water at a temperature of 100° C.?

Substituting in the formula we have, $P = \frac{373}{1} \times \frac{22.4}{273}$. This would give the pressure due to one gram-molecule of sugar. Since only 114 g. were used the pressure would be $\frac{114}{342}$ as great.

The osmotic pressures produced by very dilute solutions of binary salts and strong binary acids are found to be nearly twice as great as the pressure produced by the corresponding molecular quantity of sugar. This is due to the fact that these substances are almost completely dissociated in very dilute solution. **118.** This results in the production of twice as many particles in the solution and therefore twice the pressure.

Exercise.

117. Compute the osmotic pressures in the following cases:

1. 10 g. of sugar in one liter of water at a temperature of 0°C.
2. 10 g. of sugar in one liter of water at a temperature of 273° C.
3. Hydrochloric acid is 90% dissociated in a solution containing .1 gram-molecule to the liter. What would be the osmotic pressure produced by 3.65 g. of hydrochloric acid in one liter of water at a temperature of 0° C.?
4. Assuming sodium hydroxide to be 90% dissociated in a solution containing .1 gram-molecule to the liter, what would the osmotic pressure due to a solution of 4 g. in one liter of water at a temperature of 100° C.?

IONS AND IONIZATION.

118. Many substances, on being dissolved in water or in other solvents, are dissociated, split up, into at least two parts called *ions*.

These ions are the same in many cases as what we have already called radicals, except that they carry charges of electricity. The positive radical carrying a positive charge is known as a positive ion, and the negative radical carrying a negative charge as the negative ion.

If a molecule contains several positive radicals and one or more negative radicals, it splits up at first into one positive ion, consisting of one positive radical and one complex negative ion. On extreme dilution it would split up into as many ions as it had radicals.

An ion is indicated as a radical with a plus or minus sign above.

Thus we say sodium chloride when put in a solution in water gives positive sodium ions and negative chlorine ions, represented thus:

$$NaCl = \overset{+}{Na} + \overset{-}{Cl}$$

Or

$$H_2SO_4 = \overset{+}{H} + \overset{-}{HSO_4}$$

$$H_3PO_4 = \overset{+}{H} + \overset{-}{H_2PO_4}.$$

It is understood that the total positive charge of all the positive ions is exactly equal to the total negative charge of all the negative ions.

The Degree of Dissociation.

119. All the strong acids and bases and all salts are almost completely dissociated in dilute solution.

Substances that are about 90% dissociated in a solution containing .1 gram-molecule per liter:

Hydrochloric acid, HCl.
Hydrobromic acid, HBr.
Hydroiodic acid, HI.
Chloric acid, $HClO_3$.
Bromic acid, $HBrO_3$.
Nitric acid, HNO_3.
Potassium hydroxide, KOH.
Sodium hydroxide, $NaOH$.
Barium hydroxide, $Ba(OH)_2$.
Strontium hydroxide, $Sr(OH)_2$.

Some substances that are very slightly dissociated in solution.

Water, H_2O.
Carbonic acid, H_2CO_3.
Acetic acid, $HC_2H_3O_2$.
Hydrogen sulphide, H_2S.
Hydrocyanic acid, HCN.
Ammonium hydroxide, NH_4OH.
Boric acid, H_3BO_3.
Aluminum hydroxide.
Ferric hydroxide.
Zinc hydroxide.

The Solution of Metals in Acids.

120. All metals when placed in liquids have a tendency to send off ions into the liquid. These ions carry positive charges of electricity, and the metal is left with a corresponding negative charge. The solution does not proceed to any great extent in water or in liquids that have no marked chemical action on the metal, on account of the attraction of the negatively charged metal for the positively charged ions. The tendency of a metal to send off ions may be called its **solution pressure.** This pressure varies greatly with the different metals, platinum and gold having very low solution pressures, and zinc, magnesium and sodium very high pressures.

If a piece of zinc or other metal with a high solution pressure be placed in dilute sulphuric acid, it at once begins to send off ions. The hydrogen ions from the solution at once neutralize the negative charge on the zinc and become *atomic* or *nascent* hydrogen. These combine with each other at once and are set free as gaseous hydrogen. The reaction proceeds rapidly with extensive evolution of hydrogen until the zinc is all dissolved, the acid all used up or the solution becomes so saturated with zinc ions that their osmotic pressure is equal to the solution pressure of the metal.

The Precipitation of Metals from Solution.

121. If a metal having a high solution pressure be placed in a solution of a salt of a metal having a low solution pressure, the latter will be deposited in the metallic form on the metal of high solution pressure. The amount deposited will be the exact chemical equivalent of the amount of the metal that goes into solution.

Suppose a large piece of copper be placed in a solution of silver nitrate. The copper, having a higher solution pressure than the silver, will send ions into solution; the silver ions will go to the copper plate, where they will give up their positive charges and appear as metallic silver. If sufficient copper is present, all the silver will be thrown out of solution. The amount will be two atoms of silver for each atom of copper dissolved.

The Simple Galvanic Cell.

122. If two different metals be placed in the same liquid they each send off ions. Each becomes negatively charged. If they be connected outside of the liquid by a wire, a current will be found to flow through the wire from the one having the lower solution pressure to the one having the higher. This is because the one having the higher pressure constantly sends positive ions into the solution, and hydrogen ions, or some other positive ions from the solution, are carried to the other metal, where they give up their positive charges and are set free.

The Daniell Cell.

123. In this cell we have a zinc plate surrounded by a solution of zinc sulphate in one part of the jar and a copper plate surrounded by a saturated solution of copper sulphate in another part. The two solutions are kept separate either by the force of gravity or by a porous cup. When the two plates are connected by a wire, the zinc, having the greater solution pressure, sends off zinc ions into the solution, thereby forcing the copper ions to the copper plate, where they give up their electrical charge and appear as a deposit of metallic copper.

PASSAGE OF THE ELECTRIC CURRENT THROUGH LIQUIDS.

124. **Electrolyte,** any substance which, in the melted or dissolved condition, will allow the passage of the electric current.

Electrodes, the terminals immersed in a liquid.

The **positive electrode** is the one by means of which the current enters the liquid, called the *anode.*

The **negative electrode** is the one by means of which the current leaves the liquid, called the *cathode.*

Only Liquids that contain Ions will allow the passage of the Current.

Pure water is almost entirely undissociated, and therefore acts like a non-conductor. If, however, any substance that is highly dissociated in water solution be added, even in very small quantity, the solution at once becomes a good conductor.

Let fig. 19 represent a vessel containing a dilute solution of sodium chloride, and let A be the positive electrode and B the negative electrode; then,

1. A has a positive charge of electricity on it and B a negative charge.

2. When the salt was dissolved, it was dissociated into positive sodium ions and negative chlorine ions.

3. Since unlike kinds of electricity attract each other, the chlorine ions, being negatively charged, move toward the positive electrode and there give up their negative charges and become atomic or *nascent* chlorine. These soon unite with each other, forming molecular or gaseous chlorine, which is set free. In like manner the sodium ions go to the negative electrode and give up their charges; but the sodium enters into reaction with the water forming sodium hydroxide, and hydrogen is set free.

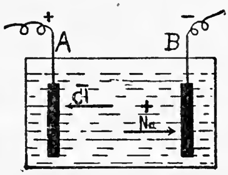

Fig. 19.

If the positive ion is a metal that does not react with the liquid, it is generally thrown out of solution in the form of a deposit on the electrode.

The negative ion, when it is a single atom of a non-metal, is generally set free; if it is a complex oxygen ion it usually reacts with the water to form an acid, and oxygen is set free.

Thus,

$$Cu(NO_3)_2 \text{ gives the ions } \overset{+}{Cu} \text{ and } \overline{NO_3}+\overline{NO_3}.$$

The copper, since it does not react with water, is disposed in the metallic form as a copper plating on the negative electrode, and the nitrate reacts with the water according to the equation,

$$4NO_3 + 2H_2O = 4HNO_3 + O_2.$$

Or we might assume that the complex oxygen ion, on losing its negative charge, breaks up into the atomic or nascent con-

dition. These atoms being very active chemically, react with each other and with the water to form nitric acid and oxygen.

Exercise.

Write all the equations showing the changes involved in the electrolysis of the following substances:

Hydrochloric acid, nitric acid, sulphuric acid, sodium chloride, sodium nitrate, sodium sulphate, sodium hydroxide. cupric chloride, cupric nitrate, cupric sulphate, silver nitrate and silver sulphate.

125. **Faraday's Law.** If any number of electrolytic cells each containing different electrolytes be connected in series and a current passed through them all, the actual quantities of the various elements set free will be proportional to the corresponding chemical equivalents of those elements.

Let us assume that the elements so set free were oxygen, hydrogen, copper, silver, lead, zinc, mercury and chlorine, and that 16 weights of oxygen were formed in the first cell.

Then there would have been formed at the same time, 2 weights of hydrogen, 63.6 of copper, 216 of silver, 206.9 of lead, 65.4 of zinc, 200 of mercury and 71 of chlorine. In other words, **the quantity of each element set free is proportional to its atomic weight divided by its valence.**

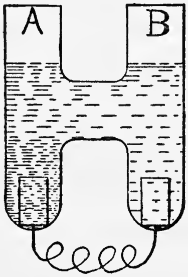

Fig. 20.

A simple Concentration Cell.

126. In general, a current will be produced, if two electrodes of the same metal are so arranged in a cell that one is surrounded with a strong and the other with a weak solution of a salt of that metal.

Let fig. 20 represent such an arrangement, the side A containing a silver plate surrounded by a strong solution of silver

nitrate and the side B a silver plate surrounded by a weak solution. On the side A, there is a greater pressure of silver ions against the silver plate than there is on the side B. No great quantity of silver can be deposited, however, on account of the inability of the positive charge to leave the plate. If the two plates be connected by a wire, the excess of positive charge from the side A is carried over to the side B and more silver is deposited on the plate A. The plate B sends silver ions into solution. A current will continue to flow through the wire until the concentration of the silver nitrate is the same on both sides.

THERMOCHEMISTRY.

127. **Law of Dulong and Petit.** The product of the specific heat and the atomic weight of all elements in the solid state is a constant and is equal to about 6.4.

In order to find the approximate atomic weight of an element which exists in the solid state, therefore, it is only necessary to divide 6.4 by its specific heat. For example, the specific heat of bismuth is .0308; its atomic weight would be $\frac{6.4}{.0308} = 208.$ 98.

128. Heat of Formation and Heat of Decomposition.

Calorie, the amount of heat necessary to raise the temperature of one kilogram of water from 0° C. to 1° C. This is known as the large calorie. A small calorie is one one-thousandth of a large calorie. In what follows, the numbers refer to the large calorie.

Heat of formation is the number of calories that are set free or absorbed when one gram-molecule of a substance is formed from the elements.

Heat of decomposition is the number of calories that are set free or absorbed when one gram-molecule of a compound is decomposed into its elements.

$$H_2 + Cl_2 = 2HCl \quad + 22 \text{ Cal.}$$
$$2K + S = K_2S \quad + 101 \text{ Cal.}$$
$$2H_2 + O_2 = 2H_2O \quad + 68 \text{ Cal.}$$
$$H_2 + S + 2O_2 = H_2SO_4 \quad + 190 \text{ Cal.}$$
$$2Na + S + 2O_2 = Na_2SO_4 + 328 \text{ Cal.}$$

129. Heat of reaction is the number of calories set free or absorbed when molecular quantities of substances enter into reaction.

An **exothermic** reaction is one in which the heat of reaction is positive, or in which heat is set free; an **endothermic** reaction is one in which the heat of formation is negative, or in which heat is absorbed.

$$NaOH + HCl = NaCl + H_2O + 13.7 \text{ Cal.}$$
$$2N_2 + O_2 = 2N_2O - 17.4 \text{ Cal.}$$

The heat of reaction is equal to the difference between the algebraic sum of the heats of formation of the factors and the heats of formation of the products.

The heat of formation will vary according to state in which the substance exists in the reaction. Table No. VII gives the heats of formation of some common compounds in the various states in which they may be formed.

Exercise.

130. Making use of the values in table VII, work out the heats of reaction in the following cases:

Sodium oxide, solid, with sulphur trioxide, solid.

Nitrogen pentoxide, solid, with sodium oxide, solid.

Sodium hydroxide, dissolved, with hydrochloric acid, dissolved.

Zinc with sulphuric acid, dissolved.

Example. What will be the heat of reaction if potassium is dissolved in hydrochloric acid?

$$\text{K (solid)} + \underset{39.3 \text{ Cal.}}{\text{HCl (dissolved)}} = \underset{101.2 \text{ Cal.}}{\text{KCl (dissolved)}} + \text{H (gas)}.$$
$$101.2 - 39.3 = 61.9 \text{ Cal.}$$

CHEMICAL EQUILIBRIUM AND THE LAW OF MASS ACTION.

Chemical Equilibrium.

131. Heretofore in this book all chemical reactions have been represented by definite chemical equations, in which very definite quantities of substances are represented as entering into chemical reaction with each other and producing very definite quantities of other substances. If the reaction be imagined to take place in an inclosed space of such a nature that none of the products can escape, it will never go on to completion as indicated by the equation, but will proceed until a definite state of equilibrium is established between the factors and the products. The quantities of the various substances in the final state will vary with the temperature, the chemical affinities of the reacting substances and the concentration of these substances that were originally present.

In other words, equilibrium will be established when the velocity of the factors forming the products is just equaled by the velocity of the products reforming the factors.

In this sense all reactions are reversible.

If A and B enter into reaction to produce two other substances C and D, then there will be in the final condition not only some of B and D but also some of A and B.

If it is desired to represent the possibility of a reversed reaction taking place, it is customary to substitute a double arrow for the equality sign, thus,

$$A+B \rightleftarrows C+D.$$

Examples:

$$C+H_2O \rightleftarrows CO+H_2$$

$$Na_2SO_4+2HCl \rightleftarrows 2NaCl+H_2SO_4$$

$$H_2S+ZnCl_2 \rightleftarrows ZnS+2HCl.$$

Completed Reactions.

132. If any of the products of a reaction be allowed to escape from the system, or if any are formed that are incapable of producing a reversed reaction, the original action goes on endeavoring to establish the required equilibrium until the original reaction is virtually complete.

Example 1. Precipitates. If one of the compounds formed in solution is insoluble, it will separate at once as a precipitate, the equilibrium will be destroyed, and more of the factors will combine until one of them is entirely used up.

$$BaCl_2 + K_2SO_4 = \underline{BaSO_4} + 2KCl.$$
$$CaCl_2 + Na_2CO_3 = \underline{CaCO_3} + 2NaCl.$$
$$Cu(NO_3)_2 + H_2S = \underline{CuS} + 2HNO_3.$$

Example 2. One of the products may be volatile at the temperature at which the reaction takes place, in which case it will escape in the form of a gas and allow the reaction to proceed to completion.

$$Na_2CO_3 + 2HCl = H_2O + CO_2 + 2NaCl.$$
$$FeS + H_2SO_4 = FeSO_4 + H_2S.$$
$$KCN + HCl = KCl + HCN.$$
$$Na_2SO_3 + 2HCl = 2NaCl + SO_2 + H_2O.$$

Example 3. One of the products may be water. In this case the water may or may not produce a reversed reaction. In most of the cases of neutralization of an acid by a base, the water formed cannot reproduce the acid and base by its action on the salt.

$$NaOH + HCl = NaCl + HOH.$$
$$H_2SO_4 + 2NaOH = Na_2SO_4 + 2HOH.$$

THE LAW OF MASS ACTION.

133. The above examples are all special cases illustrating the **Law of Mass Action.** According to this law, **when a reaction has reached a condition of equilibrium, the ratio of the products of the concentrations of the substances on the two sides of the equation is always equal to a constant, no matter what the concentration was at the start.** This is strictly true only of equations containing one molecule of a kind.

By concentration is meant the number of molecular weights per liter.

A substance, A, reacts with B to form C and D. Let **a**, **b**, **c** and **d** represent the resulting concentrations. The equation representing the reaction would be,

$$\underset{\mathbf{a}}{A} + \underset{\mathbf{b}}{B} \rightleftarrows \underset{\mathbf{c}}{C} + \underset{\mathbf{d}}{D}.$$

According to the law then,

$$\frac{ab}{cd} = k,$$

a constant. This constant will vary with the chemical affinities involved and with the temperature.

If the equation indicated two or more molecules of any one of the substances the concentration would have to be raised to the corresponding power. Thus if there were two molecules of A, the expression would become,

$$\frac{a^2b}{cd} = k.$$

The law admits of almost universal application to chemical reactions as well as physical. The following examples illustrate its application.

Example 1. **Solution of Gases in Liquids.**

134. Consider the case of the solution of nitrogen in water.

$$N_2 \text{ gas} \rightleftarrows N_2 \text{ dissolved.}$$

When a condition of equilibrium has been reached, the ratio of the concentration of the nitrogen in the gaseous form to the concentration in the dissolved form must be a constant.

What would be the effect of increasing the pressure on the gas?

That is, in the expression

$$\frac{a}{b} = k,$$

if we increase **a**, what must happen to **b** in order that **k** remain the same? Evidently **b** must increase in proportion. In other words, if the pressure of a gas in contact with a liquid be increased, more of the gas will dissolve in the liquid. The following law is a direct consequence of the law of Mass Action:

Henry's Law. *At a given temperature the amount of gas that dissolves in a liquid is proportional to the pressure.* This does not hold for those gases that are extremely soluble, such as ammonia, hydrochloric acid, etc.

Example 2. **A Liquid in equilibrium with its Vapor.**

135.

$$\underset{a}{H_2O \text{ vapor}} \rightleftarrows \underset{b}{H_2O \text{ liquid.}}$$

$$\frac{a}{b} = k.$$

Since the concentration of the water cannot be changed, the concentration of the vapor cannot be altered. Any attempt to increase it results in transforming some of it to the liquid condition.

Example 3. **A solid in contact with its solution.**

136. $$\underset{\mathbf{a}}{\text{NaCl dissolved}} \rightleftarrows \underset{\mathbf{b}}{\text{NaCl solid.}}$$

Since the concentration of the solid cannot be changed, **b** is a constant, therefore **a** cannot be changed. Any attempt to increase the concentration of **a**, by reducing the volume of the liquid, results in causing some of the salt to crystallize.

Example 4. **Dissociation of Ammonium Chloride.**

137. If ammonium chloride be heated, it is broken up into ammonia and hydrochloric acid.

$$\underset{\mathbf{a}}{NH_4Cl} \rightleftarrows \underset{\mathbf{b}}{NH_3} + \underset{\mathbf{c}}{HCl}.$$

When a condition of equilibrium is reached we have,

$$\frac{\mathbf{a}}{\mathbf{bc}} = \mathbf{k}.$$

What would be the effect of having a large quantity of ammonia present? Evidently **b** would be much larger and **c** would have to become much smaller in order that **k** remain unchanged. But if there were very little HCl formed there must be very little ammonium chloride dissociated. The presence of a large amount of ammonia prevents the dissociation of ammonium chloride by heat.

Example 5. **Ionization.**

138. When an acid, base or salt is dissolved it is ionized to a greater or less extent. A definite equilibrium is established between the undissociated material and the ions, and the law of Mass Action holds true. Sodium chloride breaks up in solution into sodium and chlorine ions, and we have,

$$\underset{\mathbf{a}}{\text{NaCl undissociated in sol.}} \rightleftarrows \underset{\mathbf{b}}{\overset{-}{Cl}} + \underset{\mathbf{c}}{\overset{+}{Na}}.$$

$$\frac{\mathbf{a}}{\mathbf{bc}} = \mathbf{k}.$$

Suppose a large quantity of hydrochloric acid be added.

This acid is highly dissociated in solution and the concentration of the Cl ions will be increased. That means that the concentration of the Na ions must be decreased or, what is the same thing, the concentration of the undissociated sodium chloride must be increased.

If the solution had been saturated, some more salt would have to crystallize out to restore the equilibrium.

The Formation of Precipitates.

139. In the case of a so-called insoluble substance surrounded by a liquid, we must recognize that some of it does go into solution, and that there is a definite equilibrium between the solid and dissolved portions. In other words, when a precipitate is formed the liquid is a saturated solution of that substance.

The undissociated dissolved substance is in equilibrium with its ions. Calling the corresponding concentrations, **a**, **b** and **c**, we have

$$\frac{a}{bc} = k.$$

Since **a** is the concentration of a substance in a saturated solution it is a constant; therefore the product of the concentrations of the ions, **bc**, in such a solution is a constant. The product of the concentrations of the ions of a substance in its *saturated solution* is known as the **Solubility Product** of that substance.

If a substance be added that increases the concentration of the ions, the product of these concentrations becomes greater than the solubility product, and more of the substance separates as a precipitate. If a substance be added that lessens the concentrations of the ions, more of the solid will go into solution.

For example, consider the case of barium sulphate.

$$\underset{a}{BaSO_4 \text{ undissociated}} \rightleftarrows \underset{b}{\overset{++}{Ba}} + \underset{c}{\overset{--}{SO_4}}.$$

$$\frac{a}{bc} = k.$$

Since this is the case of a saturated solution, **bc** is the solubility product and is a constant.

If an excess of barium chloride be added to a solution of sodium sulphate, the solution would for an instant contain very high concentrations of barium and sulphate ions; the product of these concentrations would be much greater than the solubility product, **bc**, and barium sulphate would be precipitated.

Precipitation of a soluble Salt from Solution by the Addition of an Excess of a Reagent having a common Ion.

140. Consider the case of a saturated solution of sodium nitrate.

The dissolved salt, undissociated, is in equilibrium with its ions.

$$\underset{\mathbf{a}}{NaNO_3} \text{ undissociated} \rightleftarrows \underset{\mathbf{b}}{\overset{+}{Na}} + \underset{\mathbf{c}}{\overset{-}{NO_3}}.$$

and we have

$$\frac{\mathbf{a}}{\mathbf{bc}} = \mathbf{k}.$$

Since **a** is constant, **bc** is a constant and equal to the solubility product of sodium nitrate.

Now suppose a large quantity of nitric acid be added to the solution: The concentration of the NO_3 ion will be increased, and since **bc** is a constant, **b** must decrease. That means that more undissociated $NaNO_3$ must be formed. The solution, however, was already saturated, therefore salt will be precipitated.

141. Again, we might consider the case of a saturated solution of silver acetate,

$$\underset{\mathbf{a}}{AgC_2H_3O_2} \text{ undissociated} \rightleftarrows \underset{\mathbf{b}}{\overset{+}{Ag}} + \underset{\mathbf{c}}{\overset{-}{C_2H_3O_2}}.$$

$$\frac{\mathbf{a}}{\mathbf{bc}} = \mathbf{k}.$$

a is a saturated solution and therefore a constant; **bc** must be a constant. If silver nitrate be added the concentration of the silver ions will increase and the concentration of the acetate ions must decrease. This would cause a precipitation of some of the salt.

The Addition of highly dissociated Acids or Salts to solutions of weak Acids.

142. If a large amount of sodium acetate be added to a solution of acetic acid, the acid properties of the solution almost disappear.

Acetic acid is a weak acid; that is, it is only slightly dissociated in solution.

$$\underset{a}{HC_2H_3O_2} \text{ undissociated} \rightleftarrows \underset{b}{\overset{+}{H}} + \underset{c}{C_2\overset{-}{H_3}O_2}.$$

The equilibrium equation would be:

$$\frac{a}{bc} = k.$$

When the sodium acetate is added, the concentration of the acetate ions becomes very much larger and therefore the concentration of the hydrogen ions must decrease. As it was already very small, the result is, almost no hydrogen ions are left in the solution, and therefore the acid nature of the solution almost disappears.

Addition of hydrochloric acid to a solution of acetic acid causes an almost complete suppression of the acetate ions.

143. $$\underset{a}{HC_2H_3O_2} \text{ undissociated} \rightleftarrows \underset{b}{\overset{+}{H}} + \underset{c}{C_2H_3\overset{-}{O_2}}.$$

The equilibrium equation gives

$$\frac{a}{bc} = k$$

Hydrochloric acid is almost completely dissociated in solution; its addition therefore causes a large increase in the concentration of the hydrogen ions with a resulting decrease in the acetate ions. As the concentration of the acetate ions was already rather small, it is reduced to almost nothing.

If hydrochloric acid be added to a solution of hydrogen sulphide, the $\overline{S}$ ions are diminished to such an extent that the solution will not blacken silver iodide paper immersed in it.

Hydrogen sulphide is a very weak acid. It is slightly dissociated in solution according to the equation,

$$\underset{\mathbf{a}}{H_2S} \text{ undissociated} \rightleftarrows \underset{\mathbf{b}}{2\overset{+}{H}} + \underset{\mathbf{c}}{\overset{--}{S}}.$$

The equilibrium equation gives

$$\frac{\mathbf{a}}{\mathbf{b}^2\mathbf{c}} = \mathbf{k}.$$

When the HCl is added, the concentration, **b**, is very greatly increased and the concentration, **c**, correspondingly decreased. This means that as a result there are so few S ions that there is no chance for the formation of sulphides with salts of metals.

The Necessity for an excess of Reagent to produce Complete Precipitation.

144. In many cases the solubility of the precipitate is sufficiently appreciable to prevent complete precipitation if exactly the theoretical amount of the precipitant is used.

In such cases it is necessary to use an excess of the reagent.

For example, consider the case of the precipitation of calcium oxalate. We have,

$$\underset{\mathbf{a}}{CaC_2O_4} \text{ undissociated} \rightleftarrows \underset{\mathbf{b}}{\overset{++}{Ca}} + \underset{\mathbf{c}}{\overset{--}{C_2O_4}}.$$

$$\frac{\mathbf{a}}{\mathbf{bc}} = \mathbf{k}.$$

a is a constant in the saturated solution; **bc** is therefore a constant and equal to the solubility product of calcium oxalate.

Suppose an excess of ammonium oxalate be added; this substance, being a salt, is highly dissociated, the concentration of the oxalate ions is greatly increased, and consequently the concentration of the calcium ions must be decreased. This necessitates a precipitation of more calcium oxalate.

The Prevention of the Formation of Precipitates.

145. Why cannot zinc sulphide be precipitated in an acid solution?

Assume that hydrochloric acid is present in considerable quantity, so that the concentration of the H ions is large.

Hydrogen sulphide is very slightly dissociated.

$$\underset{a}{H_2S \text{ undissociated}} \rightleftarrows 2\underset{b}{\overset{+}{H}} + \underset{c}{\overset{--}{S}}.$$

$$\frac{a}{b^2c} = k.$$

The addition of the hydrochloric acid vastly increases the concentration of the hydrogen ions and consequently decreases the concentration of the S ions.

If zinc be present in the above solution, the product of the concentration of the zinc and sulphide ions is very small and consequently less than the solubility product of zinc sulphide. Therefore zinc sulphide will not precipitate, and if it were present it would be dissolved.

If copper were present in the solution, the product of the concentrations of the sulphur and copper ions would be greater than the solubility product of copper sulphide, and there would be a precipitation of the latter.

Why does Calcium Oxalate dissolve in Hydrochloric Acid?

146. The equilibrium equation gives

$$\underset{a}{CaC_2O_4 \text{ undissociated}} \rightleftarrows \underset{b}{\overset{++}{Ca}} + \underset{c}{\overset{--}{C_2O_4}}.$$

$$\frac{a}{bc} = k$$

or since a is a constant, bc is a constant. If hydrochloric acid be added, a large quantity of hydrogen ions are introduced into the solution, and since oxalic acid is only slightly dissociated, the concentration of the oxalate ions will be decreased because they form undissociated oxalic acid and more of the calcium oxalate must dissolve to make up the loss.

Action of Ammonium Chloride on Magnesium Hydroxide.

147. Ammonium hydroxide produces a precipitate of magnesium hydroxide, because the product of the concentrations of the magnesium ions and the hydroxide ions is greater than the solubility product of magnesium hydroxide.

Now, if a large quantity of ammonium ions be added in the form of ammonium chloride, the concentration of the hydroxide ions of the ammonium hydroxide must decrease. In the equilibrium equation,

$$\underset{\mathbf{a}}{Mg(OH)_2} \text{ undissociated} \rightleftarrows \underset{\mathbf{b}}{\overset{++}{Mg}}+2\underset{\mathbf{c}}{\overset{-}{OH}}.$$

$$\frac{\mathbf{a}}{\mathbf{bc}^2} = \mathbf{k}.$$

a is a constant, therefore **bc** must be a constant.

If the concentration of the OH be decreased, the product **bc** would be less than the solubility product of magnesium hydroxide. That means, that there would not only be no precipitation of magnesium hydroxide; but if there were any of the undissolved material present, it would have to go into solution to bring up the concentration, **b**, to the required point.

Hydrolysis.

148. Many salts are hydrolyzed by water. That is they react with the water to reform the acid and base from which they were originally formed. This is true of salts that are either compounds of strong acids with weak bases, strong bases with weak acids or weak acids with weak bases.

By a strong acid or base is meant one that is highly dissociated in water solution.

For example, a water solution of potassium cyanide reacts alkaline. It is a compound of the strong base, KOH, with the weak acid, HCN.

$$KCN+HOH \rightleftarrows KOH+HCN.$$

Since the KOH is highly dissociated in solution, it would furnish a large quantity of OH ions, and since the hydrocyanic acid is a weak acid and only slightly dissociated, it would not furnish enough hydrogen ions to neutralize the alkaline effect of the OH.

Zinc sulphate in water reacts acid. Sulphuric acid and zinc hydroxide are theoretically formed. The acid is highly dissociated and furnishes hydrogen ions. Zinc hydroxide is a weak

base and prevents the formation of enough OH ions to neutralize the effect of the hydrogen ions.

Bismuth trichloride is actually dissociated into bismuth oxychloride and hydrochloric acid, according to the equation,

$$BiCl_3 + H_2O = \underline{BiOCl} + 2HCl,$$

giving a precipitate of the oxy-chloride and forming free hydrochloric acid.

Exercise.

Explain the reaction of the following substances with water: Antimony chloride; zinc chloride; copper sulphate; aluminum sulphate; potassium sulphide.

PART IV.

MISCELLANY.

IMPORTANT QUESTIONS RELATING TO THE FUNDAMENTAL IDEAS OF ELEMENTARY CHEMISTRY WHICH EVERY STUDENT SHOULD BE ABLE TO ANSWER AT THE END OF FOUR MONTHS OF STUDY.

149. The numbers following the question refer to the paragraph in which the answer will be found.

1. What is matter? **1.**

2. What is an element; how many elements are there; how many of these are common? **2.**

3. What is the symbol of an element? **3.**

4. Give the symbols and names of some common elements in groups. **4.**

5. What is meant by the expression "Properties of matter"? **5.**

6. What is a molecule; an atom? **6, 7, 8.**

7. What is meant by the chemical composition of a substance? **9.**

8. What is a physical change; a chemical change? Give five examples of each. **10, 11.**

9. What is a chemical compound? What kinds of molecules would it contain? **12.**

10. What is a mechanical mixture? What kinds of molecules would it contain? **13.**

11. What is an atomic weight? **14.**

12. What is a molecular weight, and how would that corresponding to a particular formula be found? **15.**

13. Give the full meaning of the symbol of an element. **16.**

14. What is the symbol or formula of a compound? **17.**

15. Give the symbols of seven common compounds. **18.**

16. Give the full meaning of the symbol or formula of a compound and illustrate from the formula of sulphuric acid. **19.**

17. What is meant by analysis? **20.**

18. How may the percentage composition of a substance be worked out from its formula? **21.**

19. How may the formula for a substance be worked out from the analysis? **23.**

20. What is a binary compound, and how is such a compound named? **26.**

21. What is the law of definite weight? **27.**

22. What is meant by the combining power of an element? **30.**

23. **How is it known that elements have different combining powers? 28-29.**

24. Give the combining powers of the different elements by groups. **31.**

25. If an element has several combining powers, which one should ordinarily be used? **31.**

26. **How do you tell what the formula of a given binary compound will be? 32.**

27. Into what two general classes may all the elements be divided? **25.**

28. Give the symbols and names of all the metals. **25.**

29. Give the symbols and names of all the non-metals. **25.**

30. Give the names of all the binary compounds of the metals with the non-metals. **26.**

31. **Give the formulas and names of all the compounds of the metals with the non-metals. 32.**

32. When are two elements said to be chemically equivalent? **34.**

33. **How many grams of any one element would be equivalent to ten grams of any other element? 34.**

34. What is a reaction? **36.** What is a reagent? **37.**

35. **What is a chemical equation? 38.**

36. When only have you a right to make use of a chemical equation? **38.**

37. Name some elements that are known to have two atoms in the molecule. **38.**

38. Write equations showing the formation of binary compounds from the elements. **39.**

39. State the full meaning of an equation, making use of any equation whatever. **41.**

40. How many grams of a binary compound may be formed from ten grams of one of the elements? **41.**

41. How many grams of any one element will it take to form ten grams of a corresponding binary compound? **41.**

42. What is a ternary compound? **43.**

43. What is a radical; a positive radical; a negative radical? **45, 46, 47.**

44. Give the symbols and names of all the radicals in groups. **50.**

45. Some radicals have two combining powers; how is a distinction made in the name? **51.**

46. Give the names and formulas of the compounds of all the positive radicals with all the negative radicals. **52, 54.**

47. What is an acid? Give ten examples, symbols and names. **50.**

48. What is a base? Give ten examples, symbols and names. **50.**

49. What is a salt? Give ten examples, symbols and names. **50.**

50. What is an alkali? Give three examples. **57.**

51. What is a basic oxide? Give five examples. **60.**

52. What is an acidic oxide? Give five examples. **61.**

53. What is an alkaline oxide? Give five examples. **62.**

54. What is an acid anhydride? Give five examples. **63.**

55. State seven general methods for the preparation of salts, tell what else is formed and discuss each case. **64, a, b, c, d, e, f, g.**

56. What is an oxidizing agent? **66.**

57. What is reduction? **67.**

58. Suppose a substance contains carbon, hydrogen, sulphur and phosphorus; what products will be formed if it is completely burned? **68.**

59. Write equations showing what takes place when some of the common elements are burned. **67.**

60. Write an equation showing the burning of alcohol, C_2H_5OH. **68.**

61. When an element forms several different oxides how are they named? **32.**

62. Mention three substances that give up oxygen when heated. **71.**

63. Mention three substances that decompose water with the formation of oxygen. **71.**

64. How can you tell whether a compound containing oxygen and two other elements is an oxidizing agent or not? **71.**

65. State the formulas of ten different oxidizing agents and indicate by equations how they oxidize. **78.**

66. **Write an equation showing the oxidation of phosphorus by nitric acid. 73.**

67. What is the action of manganese dioxide on hydrogen chloride, hydrogen bromide and hydrogen iodide? **75.**

68. Indicate the oxidation of hydrogen chloride with potassium permanganate. **76.**

69. Show by an equation how to change cuprous chloride to cupric chloride by means of nitric acid. **77.**

70. Indicate the oxidation of ferrous sulphate to ferric sulphate by means of potassium permanganate. **77.**

71. Show by equations the oxidation of each of ten reducing agents by each of ten oxidizing agents. **78, 79.**

72. Mention ten good reducing agents. **80.**

73. Show by equations the reduction of ten metallic oxides with three different reducing agents. **80, 81, 82.**

74. What is oxidation? **65.**

75. Write equations showing the formation of any salt in six different ways. **64.**

Important Questions relating to the more advanced Theory.

150. 1. State Avogadro's Hypothesis. **84.**

2. How is it known that the hydrogen molecule is divisible? **86.**

3. How is it known that the molecule of hydrochloric acid contains equal numbers of hydrogen and chlorine atoms? **87.**

4. How is it known that the hydrochloric acid molecule contains but one hydrogen atom? **88.**

5. How is it known that the hydrogen molecule contains two atoms? **89.**

6. What is gas density? **90.**

7. What is vapor density? **91.**

8. Why is the molecular weight equal to twice the gas density? **92.**

9. Describe two methods for getting gas density. **93, 94.**

10. How is a formula determined from the analysis and the gas density? **95.**

11. Determine the simplest formula from the analysis. **95.**

12. How are atomic weights determined? **98.**

13. Describe the nature of gaseous bodies. **99.**

14. How does the velocity of a gas molecule vary with the gas density? **99.**

15. State Boyle's Law. **100.**

16. How does the volume of a body of gas vary with the temperature? **101.**

17. State Charles' Law. **102.**

18. State the general gas law. **103.**

19. Given the volume of a body of gas at definite temperature and pressure; find the volume, pressure or temperature under other conditions. **104.**

20. What is meant by Normal Temperature and Pressure? **105.**

21. Describe the nature of a vapor. **107.**

22. What is meant by the Boiling Point? **107.**

23. Under what conditions will a vapor condense? **108.**

24. Define the terms Critical Temperature and Pressure. **109.**

25. State the Boiling Point Law. **111.**

26. State the Freezing Point Law. **112.**

27. What is meant by osmotic pressure? **114.**

28. State the laws of osmotic pressure. **115.**

29. What would be the osmotic pressure of a known weight of a substance in a given volume of solvent at a given temperature? **116.**

30. What are Ions? **118.**

31. Mention ten substances that are highly dissociated in solution. **119.**

32. Mention ten substances that are very slightly dissociated in solution. **119.**

33. Explain the solution of a metal in an acid. **120.**

34. Explain the precipitation of metals from solution. **121.**

35. Explain the action of a simple galvanic cell. **122.**

36. Explain the passage of the electric current through liquids. **124.**

37. Explain the electrolysis of potassium sulphate solution. **124.**

38. State Faraday's Law. **125.**

39. Describe a simple concentration cell. **126.**

40. State the Law of Dulong and Petit. **127.**

41. What is meant by the Heat of Formation and Heat of Decomposition? **128.**

42. What is meant by the Heat of Reaction? **129.**

43. What is meant by an Exothermic reaction? **129.**

44. What is meant by an Endothermic reaction? **129.**

45. Derive the heat of reaction in a simple case, making use of Table VII. **130.**

46. What is meant by Chemical Equilibrium? **131.**

47. In what cases do reactions go on to completion? **132.**

48. State the Law of Mass Action. **133.**

49. State Henry's Law. **134.**

50. Explain why ammonium chloride is not dissociated by heat in the presence of a large quantity of ammonia. **137.**

51. What is meant by the term Solubility Product? **139.**

52. What are the conditions necessary for the formation of a precipitate? **139.**

53. What is the effect of adding a highly dissociated acid or salt to a solution of a weak acid? **143.**

54. Explain the prevention of the formation of a precipitate of magnesium hydroxide by means of ammonium chloride. **147.**

55. What is meant by Hydrolysis? Give several examples. **148.**

College Examination Questions.

1. Describe two modifications of oxygen and state how they may be prepared.

2. Mention five substances that give up oxygen when heated.

3. Describe a test for ozone which distinguishes it from bromine or chlorine.

4. How may hydrogen peroxide be prepared and describe some of its properties?

5. How could you recognize the products that are formed when sulphur, carbon, iron, phosphorus and hydrogen are burned in oxygen?

6. Describe your laboratory method for the preparation of hydrogen and mention all precautions necessary.

7. How may oxygen and nitrogen be caused to combine?

8. How could you prove that the manganese dioxide undergoes no change when used with potassium chlorate in the preparation of oxygen?

9. What are the constituents of the atmosphere? Give the approximate proportions of the four most abundant constituents.

10. What oxides does nitrogen form and which of these are anhydrides?

11. State the exact meaning of the symbols, H; S; Fe. **16.**

12. State the exact meaning of the formulas, H_2; S_8; H_2SO_4. **19.**

13. State the full meaning of the equation, $Zn + H_2SO_4 = ZnSO_4 + H_2$. **41.**

14. State seven general methods for the preparation of salts. **64.**

15. Write equations showing the formation of sodium sulphate in six different ways. **64.**

16. What is a diabasic acid; a diacid base; a basic salt; an acid salt? Lab. Ex. No. 25.

17. How many liters of oxygen will 13 grams of mercuric oxide yield? **152.**

18. How many grams of sulphuric acid would be necessary to form 10 liters of hydrogen? **152.**

19. What are the halogens? Describe each briefly.

20. What kind of a chemical agent is required to free chlorine from hydrochloric acid? **75, 76.**

21. Explain under what conditions, if any, hydrochloric acid acts upon the following substances and what is formed in each case: Sodium hydroxide, zinc oxide, zinc, iron sulphide, sodium carbonate, barium dioxide, manganese dioxide. Write all equations.

22. What is the test for the chlorine radical?

23. How could you change sulphur into hydrogen sulphide?

24. How did you make sulphuric acid and how did you recognize the product?

25. Under what condition, if any, does sulphuric acid act upon the following substances: Copper, lead, iron, carbon, sulphur, sodium carbonate, barium dioxide, calcium fluoride, sodium chloride? **17.**

26. State two general methods for the preparation of the halogen acids.

27. Explain the use of hydrogen sulphide in the separation of copper and zinc.

28. How did you make sodium sulphite from sulphur?

29. State exactly what takes place when sodium hydroxide is neutralized with hydrochloric acid. **118.**

30. What is the effect of mixing hydrogen sulphide water and nitric acid? **79.**

31. Describe the changes that sulphur undergoes when slowly heated.

32. State three different methods for the preparation of sulphur dioxide. **71, 65, 64 d.**

33. How would you change sodium sulphate to sodium chloride? **64 e.**

34. State two methods for the preparation of nitrogen.

35. Give the equation for the action of ammonium hydroxide on hydrochloric acid. **64 c.**

36. What would be the effect of heating ammonium chloride very hot?

37. What is the action of sulphuric acid on ammonium hydroxide? **64 c.**

38. How did you make ammonium nitrate and how did you tell when it was sufficiently evaporated? **64 c.**

39. What actually happens when ammonium chloride is heated?

40. How may ammonia be obtained from an ammonium compound?

41. What is the effect of heating a mixture of ammonium sulphate and common salt?

42. How did you make nitric acid and what is its chief characteristic?

43. Write the equations showing the action of nitric acid on copper.

44. If copper scale is dissolved in dilute nitric acid, what compound would be formed? **71.**

45. Suppose sodium hydroxide were added to this solution, and the solution heated, what would be the precipitate be?

46. Suppose copper scale were dissolved in dilute hydrochloric acid, what compound would be formed?

47. If sodium hydroxide were added to this solution and the solution heated, how would this precipitate compare with the last?

48. What is aqua regia? Represent its action by equations.

49. What is the action of strong nitric acid on zinc?

50. What is the action of very dilute nitric acid on zinc? Give equations.

51. What is the action of an excess of strong nitric acid on yellow ammonium sulphide? **71.**

52. How may red phosphorus be changed to yellow and yellow to red?

53. What is the effect of adding bromine water to red phosphorus? **71.**

54. What is the effect of adding iodine water to red phosphorus? **71.**

55. What is the effect of dissolving red phosphorus in nitric acid, adding silver nitrate and ammonia so as not to mix?

56. Describe the effect of heating arsenic in closed tube.

57. Describe the effect of heating arsenic in air.

58. What is the effect of heating arsenious oxide with charcoal? **80.**

59. What is the effect of heating arsenious oxide in closed tube?

60. What effect has bromine water on metallic arsenic? **71.**

61. What is the color of silver arsenate?

62. What is formed when arsenic is dissolved in concentrated nitric acid? **73.**

63. What is the reaction between ammonium hydroxide and arsenic trioxide? What is the color of silver arsenite? **64 g.**

64. Which is the most soluble, arsenious oxide or arsenic oxide?

65. What is formed when bromine water acts on arsenious oxide? **71.**

66. What is the action of concentrated nitric acid on powdered antimony? **73.**

67. What is formed when bismuth carbonate is dissolved in hydrochloric acid? **64 d.**

68. When this solution is poured into an excess of water, what is the precipitate? **148.**

69. What are the three allotropic forms of carbon and what are your reasons for believing them to be the same substance?

70. What is the effect of heating copper scale with charcoal? **80.**

71. What is the effect of passing illuminating gas over hot cupric oxide? **80.**

72. What effects are produced by passing an excess of carbon dioxide into lime water? **64 g.**

73. What is the effect of heating hydrogen sodium carbonate?

74. What acid does carbon dioxide form with water? **69.**

75. Explain the action of a water solution of sodium carbonate on litmus. **148.**

76. What is common baking powder? Lab. Ex. No. 25.

77. What is the effect of heating Rochelle Salts until all carbon is burned?

78. How did you make orthosilicic acid? Lab. Ex. No. 26, 3.

79. Mention three ways of dissolving sand.

80. How did you make stannous chloride? **64 a.**

81. How would you change stannous chloride into stannic chloride? **77.**

82. What is the action of stannous chloride upon mercuric chloride? Lab. Ex. No. 27, 2.

83. What is the action of hydrogen sulphide upon stannous chloride; upon stannic chloride?

84. How do the common acids affect lead?

85. What is the action of a soluble sulphate upon a solution of a lead salt?

86. What is the effect of trying to dissolve red lead in dilute nitric acid? What is supposed to be the composition of red lead?

87. What is the Solvay or Ammonia process of making sodium carbonate?

88. How could you distinguish between carbonate and bicarbonate of soda?

89. Describe how you made sodium hydroxide in the laboratory. Lab. Ex. No. 29.

90. What is the effect of adding concentrated hydrochloric to a saturated solution of common salt? **140.**

91. What effect has silver nitrate upon a solution of potassium chlorate?

92. Describe methods for making the following salts: Sodium sulphate from sodium chloride, sodium chloride from sodium sulphate, sodium acetate from sodium chloride, sodium chloride from hydrogen di-sodium phosphate.

93. What oxides does copper form? Describe each.

94. Describe the action of the common acids on copper, both hot and cold. Give equations where there is any action.

95. How could you change cupric chloride to cuprous chloride? **82.**

96. How could you change cuprous chloride to cupric chloride? **73.**

97. What is the effect of adding potassium iodide to copper sulphate solution?

98. Describe all changes when sodium hydroxide is added to a solution of cupric sulphate and the mixture heated to boiling.

99. State two methods for making cuprous oxide. Lab. Ex. No. 31, 3, 5.

100. What is the effect of adding a piece of iron wire to a copper sulphate solution? **121.**

101. How did you obtain pure silver from a silver coin?

102. How could you obtain silver nitrate from silver chloride?

103. How can the precipitation of magnesium hydroxide by ammonia be prevented without making the solution acid? **147.**

104. Describe lime, slaked lime, lime water, calcium hydroxide, air-slaked lime.

105. How could you distinguish air-slaked lime from dry-slaked lime?

106. What is the difference between chloride of lime and calcium chloride? State the uses of each.

107. What is the formula of gypsum? What is plaster of Paris?

108. Why does plaster of Paris set?

109. Describe barium hydroxide. What happens when a solution of barium hydroxide is left exposed to the air? **46 g.**

110. What is the action of zinc on potassium hydroxide solution?

111. What is formed when potassium hydroxide is added to a zinc sulphate solution? Lab. Ex. No. 30, 7.

112. What is the action of zinc chloride on litmus and explain the result? **148.**

113. What is the action of cold and hot nitric acid on mercury?

114. What is the effect of adding a sodium hydroxide solution to a solution of mercuric nitrate?

115. What is the effect of adding a soluble chloride to a mercurous solution?

116. What is the effect of adding sodium hydroxide to a solution of a mercuric salt?

117. What is the action of the common acids on aluminum?

118. What is the residue after heating aluminum chloride very hot?

119. What peculiar property has aluminum amalgam?

120. Describe chromic hydroxide.

121. How could you change chromic sulphate to a chromate?

122. How could you change chromium in potassium dichromate to chromic sulphate?

123. What would be the effect of chromic anhydride on hydrochloric acid?

124. What is the effect of adding oxalic acid to potassium permanganate solution in the presence of sulphuric acid? **76.**

125. Describe the effects of the common acids on iron.

126. Describe ferrous hydroxide.

127. How did you change ferrous sulphate to ferric sulphate? **77.**

128. Describe ferric hydroxide.

129. How would you change ferric hydroxide to ferric chloride and this to ferrous chloride? **64 c, 82.**

130. Describe the effect of adding hydrogen sulphide to both ferrous and ferric salts in solution. **81.**

131. How did you make ferrous ammonium sulphate?

132. How could you change potassium ferrocyanide to ferric chloride?

133. Why does not ammonium hydroxide give a precipitate of ferric hydroxide in a solution of potassium ferrocyanide?

134. How could you make chloroform?

135. How could you make ethyl acetate?

136. What is the general formula of an alum, and how could you make ammonium aluminum alum?

137. Explain why the same amounts of heat are generated when equivalent amounts of different acids are neutralized with bases.

138. Describe the properties and preparation of the two chlorides of mercury.

139. What action, if any, has dry ammonia gas on litmus paper?

140. How is the aluminum ion detected?

141. Describe several different tests for ferrous and ferric ions.

142. What acids does phosphorus form? **69.**

143. How could you make lead chloride and what remarkable property has this salt?

144. How much methyl alcohol should be dissolved in one liter of water to give the same pressure that 5 grams of ethyl alcohol give when dissolved in the same volume of the same solvent? **115.**

145. What is meant by the term "weak acid" or "weak base"?

146. What is the method for finding the atomic weight of an element like calcium that does not form gasifiable compounds? **127.**

147. State two methods for making aluminum chloride from alum.

148. Explain the fact that the gas density of sulphuric acid is 24.5.

149. Describe and explain what takes place when a piece of zinc is placed in a solution of a lead salt. **121.**

150. Examples involving weights only. **151.**

151. Examples involving weights and volumes of gases. **152.**

152. Examples involving volumes of gases only. **154.**

153. Examples involving volumes of liquids having a certain specific gravity and a certain per cent. of material in solution. **155.**

154. Derivation of formulas from analysis and vapor density. **95.**

155. General Theory. See questions with references. **150.**

SIMPLE PROBLEMS IN ELEMENTARY INORGANIC CHEMISTRY.

151. How many grams of a substance are necessary to combine with or produce a certain number of grams of another substance?

Example. Ten grams of sodium nitrate could be formed from how many grams of nitric acid?

It is generally best to write an equation showing the possible reaction, although if the chemical equivalents are surely known this is not necessary.

$$NaOH + \underset{63}{\overset{?\ g.}{HNO_3}} = \underset{85}{\overset{10\ g.}{NaNO_3}} + H_2O.$$

From this it appears that one molecule of nitric acid is exactly equivalent to one molecule of sodium nitrate.

That is, that 63 weights of nitric acid will exactly form 85 weights of the salt. One weight would form $\frac{1}{63}$ of 85 weights, nd therefore as many weights would be necessary to form ten as $\frac{85}{63}$ is contained in 10 or $\frac{63}{85} \times 10$.

Or it might be done by proportion, which amounts to exactly the same thing. **The unknown weight in grams is to the given weight in grams as the molecular weight of the unknown substance is to the molecular weight of the known substance.**

This would give $\frac{X}{10} = \frac{63}{85}$ or $X = \frac{63}{85} \times 10$.

152. Examples involving a weight of a substance and the volume of a gas.

Example. How many liters of carbon dioxide, at normal temperature and pressure, could be formed from ten grams of sulphuric acid?

$$Na_2CO_3 + \underset{98}{\overset{10\ g.}{H_2SO_4}} = Na_2SO_4 + H_2O + \underset{44}{CO_2}.$$

One molecule of sulphuric acid is equivalent to one molecule of carbon dioxide. Ten grams of the acid would therefore form $\frac{44}{98} \times 10$ grams of carbon dioxide.

The weight of one liter of any kind of gas at normal temperature and pressure is equal to one half its molecular weight multiplied by .09.

Therefore the weight of a liter of carbon dioxide equals $22 \times .09 = 1.98$ gr. There will be as many liters in the above weight of the gas as 1.98 g. is contained in $\frac{44}{98} \times 10$ or $\frac{44 \times 10}{98 \times 1.98}$ liters at norm. temp. and pressure.

Example. How many grams of sodium sulphite would ten liters of sulphur dioxide form?

One liter of sulphur dioxide weighs $32 \times .09$ grams or 2.88 g. Ten liters would weigh 28.8 g.

$$\begin{array}{llll} SO_2 + 2NaOH = & Na_2SO_3 + H_2O. \\ 64 & 126 \\ 28.8\text{ g.} & ?\text{ g.} \end{array}$$

One molecule of the gas is equivalent to one molecule of the salt. 28.8 g. would therefore make $28.8 \times \frac{126}{64}$ g. of sodium sulphite.

153. Given the volume of a body of gas at a certain temp. and pressure to find the volume corresponding to any other temp. and pressure.

Example. A volume of nitrogen is measured at a temp. of 10° C. and pressure of 780 mm. and found to be 10 c.c. What would be the corresponding volume at 20° and 790 mm.?

The volume of a body of gas varies as its absolute temperature. In this case the temperature has increased, therefore the volume would be increased. The absolute temperatures concerned are 283° and 293°, therefore the given volume should be multiplied by $\frac{293}{283}$. This will give the new volume, considering temperature alone.

The volume of a body of gas varies inversely as the pressure. The pressure in this case has increased, therefore

the volume would be decreased on account of the pressure change. Therefore the above new volume should be multiplied by $\frac{780}{790}$.

The final reduced volume would be $10 \times \frac{293}{283} \times \frac{780}{790}$ c.c.

154. **Examples involving volumes of gases only.**

Example. What volume of oxygen would be necessary to exactly burn ten liters of alcohol vapor? What would be the volumes of the gaseous products?

$$C_2H_6O + 3O_2 = 3H_2O + 2CO_2.$$

First write a *molecular* equation representing the reaction.

Equal volumes of all gases at the same temp. and pressure contain equal numbers of molecules.

In other words, **all gaseous molecules occupy equal space.** The coefficients in a molecular equation represent at once the relative combining volumes of the indicated substances **in the gaseous form.**

That means in the above equation that one volume of alcohol vapor would combine with three volumes of oxygen and form three volumes of water vapor and two volumes of carbon dioxide. Or that ten liters of alcohol vapor would require thirty liters of oxygen to burn it, and there would be formed thirty liters of water vapor and twenty liters of carbon dioxide.

Example involving the per cent. of substance in solution and the specific gravity of the solution.

155. In the majority of chemical operations the reagents are handled in solution, the specific gravity of the liquid and the per cent. of the substance in solution being known.

The problem is exactly like the simple case in **151** except that in place of a definite weight being given or required, a volume of a certain solution, containing a certain per cent. of material and having a given specific gravity, is given or required.

If a volume of a liquid is given, find at once the amount of available material it contains and proceed as in **151.**

If the volume of a liquid is required, first find the number of grams required as in **151**, and then reduce this number to the volume of the solution necessary to give it.

It may be that the volume of one liquid is given to find the volume of another liquid concerned.

156. Example 1. 100 c.c. of sulphuric acid solution having a s.g. of 1.84 and containing 98% acid, will form how many grams of hydrochloric acid?

In all of these examples if we call S, the specific gravity of the liquid, P, the per cent. of substance in solution, and W, the weight of material in volume, V, the following expression holds true,

$$SVP = W \text{ or } V = \frac{W}{SP}, \; P = \frac{W}{SV} \text{ and } S = \frac{W}{VP}.$$

In using these expressions never confuse volumes of gases with volumes of liquid.

100 c.c. of the acid would contain $100 \times 1.84 \times .98$ grams of H_2SO_4, which equals 180.3 g.

$$\begin{array}{lcr} H_2SO_4 + 2NaCl = & Na_2SO_4 + & 2HCl. \\ 98 & & 73 \\ 180.3 \text{ g.} & & ? \text{ g.} \end{array}$$

One molecule of sulphuric acid is equivalent to two molecules of hydrochloric acid. That is, 180.3 g. of sulphuric acid would produce $\frac{73}{98} \times 180.3$ g. of hydrochloric acid.

157. Example 2. How many c.c. of a barium chloride solution, having a s.g. of 1.1 and containing 10% barium chloride, would precipitate 2 g. of sulphuric acid?

$$\begin{array}{llll} BaCl_2 + & H_2SO_4 = & \underline{BaSO_4} + & 2HCl. \\ 208 & 98 & & \\ ? \text{ g.} & 2 \text{ g.} & & \end{array}$$

$\frac{208}{98} \times 2$ g. of barium chloride would be necessary. This amount would be contained in as many c.c's. of the solution as, $SP = .11$ is contained in $W = \frac{208}{98} \times 2$ or $\frac{208 \times 2}{98 \times .11}$ c.c.

158. Example 3. How many c.c. of a solution of potassium permanganate, one c.c. of which contains .1 g. of the salt, would be necessary to oxidize all the iron contained in 100 c.c. of a solution of ferrous sulphate having a s.g. of 1.1 and containing 10% of $FeSO_4$?

SVP = W. $1.1 \times 100 \times .10 = 11$ grams of $FeSO_4$ contained in the solution to be oxidized.

$$10FeSO_4 + 2KMnO_4 + 8H_2SO_4 = 5Fe_2(SO_4)_3 + K_2SO_4 + 2MnSO_4 + 8H_2O.$$

1520	316
11 g.	? g.

The number of grams of potassium permanganate necessary would be $\frac{316}{1520} \times 11$, and this would be contained in as many c.c. of the solution as .1 is contained in it, or $\frac{316 \times 11}{1520 \times .1}$.

159. Example 4. How many c.c. of hydrochloric acid solution having a s.g. of 1.2 and containing 40% acid would be necessary to form 100 liters of carbon dioxide?

One liter of carbon dioxide weighs $\frac{44}{2} \times .09 = 1.98$ g. 100 liters would weigh 198 g.

$$CaCO_3 + 2HCl = CaCl_2 + H_2O + CO_2.$$

73	44
? g.	198 g.

The number of grams of hydrochloric acid would be $\frac{73}{44} \times 198$ g.

$V = \frac{W}{SP} = \frac{73 \times 198}{1.2 \times .40 \times 44}$ = number of cubic centimeters of the solution.

Review Exercise.

160. There are many important chemical transformations in elementary chemistry with which the student should be very familiar. Some of these are listed in Table I. The student should learn to state the various steps as actually carried out in practice, as well as to represent the successive changes by equations. It is safe to say that any student who can do this as

called for in Table I. is necessarily well grounded in the principles of elementary chemistry. If there are two or more methods he should learn them all, especially the common technical method. If the laboratory or analytical method is different from the technical method, that should also be thoroughly learned.

For example, consider No. 23, Table I. How change mercury to mercuric chloride?

Technical. Dissolve the mercury in sulphuric acid, getting mercuric sulphate. Heat the mercuric sulphate with sodium chloride and mercuric chloride will sublime out.

$$Hg + 2H_2SO_4 = HgSO_4 + SO_2 + 2H_2O.$$
$$HgSO_4 + 2NaCl = HgCl_2 + Na_2SO_4.$$

Laboratory method. Dissolve the mercury in aqua regia.

$$2HNO_3 = H_2O + 2NO + 3O.$$
$$3O + 6HCl = 3H_2O + 3Cl_2.$$
$$3Hg + 3Cl_2 = 3HgCl_2$$

Adding these equations we have,

$$3Hg + 2HNO_3 + 6HCl = 3HgCl_2 + 4H_2O + 2NO.$$

The student is not expected to invent any of the reactions, he is expected to look up the actual methods in the text book. All reactions represented must be possible.

TABLE I.

REVIEW EXERCISE.

161. Indicate by equations the following transformations.

1. Sodium to sodium chloride. **64 a.**
2. Calcium oxide to calcium nitrate. **64 b.**
3. Aluminum sulphate to aluminum hydroxide. **64 e.**
4. Sodium chloride to sodium sulphate. **64 d.**
5. Aluminum hydroxide to aluminum sulphate. **64 c.**
6. Calcium carbonate to calcium chloride. **64 d.**
7. Sodium sulphate to sodium chloride. **64 e.**
8. Sulphur trioxide to sodium sulphate. **64 f, g.**
9. Sodium chloride to sodium.
10. Calcium to calcium nitrate. **78, 8.**
11. Calcium chloride to calcium nitrate. **64 e, 64 d.**
12. Barium nitrate to barium chloride.
13. Cupric sulphate to cupric oxide.
14. Mercuric oxide to mercury.
15. Mercuric chloride to mercuric oxide.
16. Barium sulphate to barium chloride. **163.**
17. Strontium chloride to strontium oxide.
18. Silver to silver chloride. **78, 8; 64 e.**
19. Copper to cupric sulphate. **71.**
20. Copper to cupric chloride. **160.**
21. Cupric chloride to cuprous chloride. **82.**
22. Silver chloride to silver nitrate. **121, 78, 8.**
23. Mercury to mercuric chloride. **160.**
24. Mercury to mercurous chloride. **160, 82.**
25. Stannous chloride to stannic chloride. **77.**
26. Ferrous chloride to ferric chloride. **77.**
27. Cuprous chloride to cupric chloride. **77.**
28. Ferrous sulphate to ferric sulphate. **77.**
29. Ferrous chloride to ferric sulphate.
30. Mercurous chloride to mercuric sulphate.
31. Stannous chloride to stannic sulphate.
32. Sulphur to hydrogen sulphide.
33. Sulphuric acid to sulphur dioxide. **71.**
34. Nitric acid to ammonia.
35. Sodium hydroxide to sodium aluminate.
36. Chromous oxide to chromic acid.
37. Potassium dichromate to potassium chromate.
38. Iodine to iodic acid.
39. Potassium dichromate to chromic chloride.
40. Manganese dioxide to permanganic acid.

TABLE II.

TABLE OF NATURAL ELEMENTS AND COMPOUNDS FROM WHICH NEARLY ALL OTHER COMMON COMPOUNDS AND ELEMENTS MAY BE MADE.

162. **First Group.**

Sodium chloride.
Sodium nitrate.
Potassium chloride.

Second Group.

Magnesium carbonate.
Magnesium sulphate.
Calcium carbonate.
Calcium sulphate.
Strontium carbonate.
Strontium sulphate.
Barium carbonate.
Barium sulphate.
Zinc carbonate.
Zinc sulphide.
Cadmium sulphide.
Mercury.
Mercury sulphide.

Third Group.

Aluminum sulphate.
Aluminum silicate.
Borax, $Na_2B_4O_7$.

Fourth Group.

Carbon.
Carbon dioxide.
Calcium carbonate.
Silicon dioxide.
Aluminum silicate.
Tin dioxide.
Lead sulphide.

Fifth Group.

Nitrogen.
Sodium nitrate.
Calcium phosphate.
Arsenic.
Arsenic sulphide.
Antimony sulphide.
Bismuth sulphide.

Sixth Group.

Oxygen.
Hydrogen oxide.
Sulphur.
Calcium sulphate.
Sulphides.
Iron pyrite.

Seventh Group.

Calcium fluoride.
Sodium chloride.
Potassium chloride.
Sodium bromide.
Potassium iodate.
Manganese dioxide.

Eighth Group.

Ferrous carbonate.
Ferric oxide.
Nickelous carbonate.
Copper carbonate.
Copper.
Silver.
Silver sulphide.
Silver chloride.
Gold.
Platinum.

Review Exercise.

163. With a view of fixing in the student's mind the natural forms in which the common elements occur, as well as the processes by which they are transformed to the common artificial compounds, the following exercise has been designed:

Table II gives some of the natural sources of the common elements, and Table III gives the names of many common elements and compounds.

The student is assigned some element or compound from Table III, by number, and he is expected to start with natural elements or compounds from Table I and state in words what operations are necessary to produce the given substance. He is not allowed to make use of any substance whose formation from the natural substances he has not first shown. He can best represent all changes by equations, and then in his recitation translate his equations into words.

The value of this exercise cannot be overestimated. It results in fixing in the mind the relative importance of certain artificial compounds and fundamental processes in a way that no ordinary laboratory or text-book work can do.

Example. Show the artificial preparation of cupric carbonate.

Natural materials. Copper; Sulphur; Oxygen; Water; Calcium carbonate; Sodium chloride.

$$S + O_2 = SO_2.$$
$$SO_2 + O = SO_3.$$
$$SO_3 + H_2O = H_2SO_4.$$

$$2H_2SO_4 + Cu = CuSO_4 + SO_2 + 2H_2O.$$
$$H_2SO_4 + 2NaCl = Na_2SO_4 + 2HCl.$$
$$Na_2SO_4 + 2C = Na_2S + 2CO_2.$$
$$Na_2S + CaCO_3 = Na_2CO_3 + CaS.$$

$$Na_2CO_3 + CuSO_4 = \underline{CuCO_3} + Na_2SO_4.$$

Here again the student is not expected to invent any reactions, but to use his text book freely to find out what actual processes are in use.

TABLE III.

Practice Exercise.

164. Indicate by equations the formation of each of the following compounds or elements from the natural elements or compounds, as given in a previous table. Use no reagents whose formation from natural substances you have not shown.

1. Sulphuric acid.
2. Hydrochloric acid.
3. Nitric acid.
4. Calcium oxide.
5. Carbon dioxide.
6. Calcium hydroxide.
7. Sodium hydroxide.
8. Sodium carbonate.
9. Sodium bicarbonate.
10. Ammonia.
11. Hydrogen sulphide.
12. Cupric oxide.
13. Cuprous oxide.
14. Cupric sulphate.
15. Ferrous chloride.
16. Ferric hydroxide.
17. Ferrous sulphate.
18. Hydrogen.
19. Iodine.
20. Bromine.
21. Aluminum bromide.
22. Ammonium acetate.
23. Cadmium chlorate.
24. Calcium hydroxide.
25. Cupric nitrate.
26. Cuprous carbonate.
27. Magnesium sulphide.
28. Nickelous sulphate.
29. Potassium phosphate.
30. Sodium oxide.
31. Lead nitrite.
32. Barium sulphite.
33. Zinc chloride.
34. Acetic acid.
35. Arsenic acid.
36. Boric acid.
37. Hydrobromic acid.
38. Hydriodic acid.
39. Hydrofluoric acid.
40. Phosphoric acid.
41. Antimony.
42. Barium.
43. Copper.
44. Tin.
45. Lead.
46. Manganese.
47. Mercury.
48. Silver.
49. Iron.
50. Zinc.
51. Aluminum.
52. Ammonium chloride.
53. Ammonium arsenate.
54. Antimony sulphide.
55. Barium chromate.
56. Chromic sulphate.
57. Hydrogen.
58. Hydrogen peroxide.
59. Iodic acid.
60. Lead chromate.
61. Manganous sulphate.
62. Mercuric chloride.
63. Mercurous chloride.
64. Mercuric oxide.
65. Mercurous oxide.
66. Mercuric nitrate.
67. Mercurous nitrate.
68. Phosphorus.
69. Potassium arsenate.
70. Potassium dichromate.
71. Potassium bicarbonate.
72. Potassium chromate.
73. Potassium permanganate.
74. Silver nitrate.
75. Silver chloride.
76. Sodium arsenate.
77. Sodium thiosulphate.
78. Stannous sulphide.
79. Stannous chloride.
80. Stannic chloride.

TABLE IV.

165. Showing the solubility of some compounds in water. The numerator shows the number of parts soluble in 100 parts of water at 100° C., the denominator the number of parts soluble in 100 parts of water at 20° C.

	Bromide.	Carbonate.	Chloride.	Chromate.	Cyanide.	Ferrocyanide	Ferricyanide.	Hydroxide.	Iodide.	Nitrate.	Oxide.	Phosphate.	Sulphate.	Sulphide.
Al	S	---	S	---	---	---	---	I	- -	S	I	I	89.1*/36.1	---
NH_4 ..	S	S	77.3/37.2	S	S	S	S	S	S	Dec./185	---	s	103.3/75.4	S
Ba	149/104	I	72/42	I	S	S	S	50*/5	S	34.2/8.6	90+/3.4	I		Dec./Dec.
Ca	S	I	S	s/.4	Dec.	S	S	.066/.15	435+/204	S	Dec.	I	.21/.24	.27/.23
Cu′ ...	I	---	I	---	I	I	I	I	I	---	I	---	---	---
Cu″ ...	S	---	S	---	I	I	I	I	Dec.	S	I	I	203/42	I
Fe″ ...	S	I	s/s	---	---	I	I	I	S	S	I	I	42/24	I
Fe‴ ...	S	---	S	---	---	I	S	I	---	S	I	I	S	I
Pb	s/s	I	5±/.97	I	s/s	I	s/s	I	.14/.04	7/52	I	I	I	I
Mg ...	S	I	S	S	---	S	S	I	S	S	I	I	73/36	Dec.
Hg′ ...	I	---	I	s/s	---	---	---	I	I	Dec.	I	Dec./I	.3/.2	I
Hg″ ...	4/.4	---	54/7	Dec.	40/10	---	---	I	I	S	I	s/I	Dec.	I
K	102/64	147+/112	56./35.	79/62	S	100±/35±	77/50±	S	209/144	247/31	Dec.	S	26/11	S
Ag	I	I	I	I	I	I	I	I	I	1500±/227	I	I	1/.5	I
Na	120/88	45*/21	40/36	S	S	90±/30±	66/20	S	312/178	1 [illegible]/[illegible]	Dec.	S	412†/140	S
Sr	200±/99	I	102*/54	S	Dec./S	S	S	18/.68	370/179	101/71	2/.6	I	I	Dec.
Sn″ ...	S	I	S	---	---	I	---	I	s/s	---	I	I	Dec.	I
Zn	S	I	S	---	I	I	---	I	S	S	I	I	653/161	I

* Anhydrous. † At 34° C.
S Very soluble. I Insoluble.
s Slightly soluble. Dec. Decomposed by boiling.

Solubility Rules.

These rules should be memorized.

166. All the hydrogen, sodium, potassium and ammonium compounds are soluble; also all the acetates, chlorates and nitrates.

All the chlorides are soluble except those of **silver, lead** and **mercurous mercury.**

All the sulphates are soluble except those of **strontium, barium** and **lead.**

All the carbonates and phosphates are insoluble except those of **sodium, potassium** and **ammonium.**

All the hydroxides are insoluble except those of **sodium, potassium, ammonium, calcium, strontium** and **barium.**

TABLE V.

PERIODIC ARRANGEMENT OF THE ELEMENTS.

167. A few of the very rare elements are omitted.

H = 1.0075.

O	I	II	III	IV	V	VI	VII	VIII		
R	RH R_2O	RH_2 RO	RH_3 R_2O_3	RH_4 RO_2	RH_3 R_2O_5	RH_2 RO_3	RH R_2O_7	RO_4		
He 4	Li 7.03	Gl 9.1	B 11.0	C 12.0	N 14.04	O 16.00	F 19			
Ne 20	Na 23.05	Mg 24.36	Al 27.1	Si 28.4	P 31.0	S 32.06	Cl 35.45			
A 39.9	K 39.15	Ca 40.1	Sc 44.1	Ti 48.1	V 51.2	Cr 52.1	Mn 55.0	Fe 55.9	Co 59.0	Ni 58.7
	Cu 63.6	Zn 65.4	Ga 70	Ge 72.5	As 75.0	Se 79.2	Br 79.96			
Kr 81.8	Rb 85.5	Sr 87.6	Yt 89	Zr 90.6	Cb 94	Mo 96		Ru 101.7	Rh 103	Pd 106.5
	Ag 108.0	Cd 112.4	In 115	Sn 119.0	Sb 120.2	Te 127.6	I 127.0			
Xe 128	Cs 132.9	Ba 137.4	La 138.9	Ce 140.25	Ta 183	W 184		Os 191	Ir 193	Pt 194.8
	Au 197.2	Hg 200	Tl 204.1	Pb 206.9	Bi 208.5					
		Ra 225		Th 232.5		U 238.5				

169. TABLE VII.

TABLE OF HEATS OF FORMATION OF SOME COMMON COMPOUNDS AT NORMAL TEMPERATURE.

Substance.	Heat of Formation.			
	Gaseous.	Liquid.	Solid.	Dissolved.
Ammonia	+12.0	----	----	+ 20.4
Ammonia chloride	----	----	+ 75.8	+ 71.9
Ammonium nitrate	----	----	+ 88.0	+ 81.8
Barium carbonate	----	----	+280.5	----
Barium chloride	----	----	+194.7	+196.8
Barium oxide	----	----	+124.2	+158.7
Calcium carbonate	----	----	+267.7	----
Calcium hydroxide	----	----	+214.2	+214.9
Calcium oxide	----	----	+131.0	+149.5
Carbon dioxide	+97.6	----	----	+101.5
Carbon disulphide	—28.7	— 22.3	----	----
Cupric chloride	----	----	+ 51.6	+ 62.7
Cupric oxide	----	----	+ 37.2	----
Hydrochloric acid	+22.0	----	----	+ 39.3
Hydrogen sulphide	+ 2.7	----	----	+ 7.3
Lead oxide (PbO)	----	----	+ 50.3	----
Lead sulphide	----	----	+ 18.4	----
Magnesium chloride	----	----	+151.0	+186.9
Magnesium oxide	----	----	+144.0	----
Magnesium sulphate	----	----	+302.3	+329.0
Nitric acid	----	+ 41.9	----	+ 49.1
Nitrogen pentoxide	----	----	+ 13.1	+ 29.8
Potassium carbonate	----	----	+278.4	+284.9
Potassium chloride	----	----	+105.6	+101.2
Potassium hydroxide	----	----	+103.2	+116.5
Potassium nitrate	----	----	+119.5	+111.0
Potassium oxide	----	----	----	+164.6
Potassium sulphate	----	----	+344.6	+338.2
Sodium bicarbonate	----	----	+227.0	+223.7
Sodium carbonate	----	----	+269.9	+275.4
Sodium chloride	----	----	+ 97.6	+ 96.4
Sodium hydroxide	----	----	+101.9	+111.8
Sodium nitrate	----	----	+111.3	+106.3
Sodium oxide	----	----	+100.2	+155.2
Sodium sulphate	----	----	+328.4	+329.0
Sulphur trioxide	----	----	+103.3	+142.5
Sulphuric acid	----	+189.9	----	+210.9
Water	+58.0	+ 68.0	+ 69.8	----
Zinc sulphate	----	----	+230.0	+248.0

170. TABLE VIII.

Tables of Length, Weight and Volume.

10 millimeters = one centimeter.

100 centimeters = one meter.

1 meter = 39.37 inches.

1 cubic centimeter of pure water at a temperature of 4° C. weighs one gram.

1000 grams = 1 kilogram.

1 kilogram = 2.2 pounds.

1000 cubic centimeters = 1 liter.

1 liter = 1.06 quarts.

Exact Equivalents.

1 inch = 25.39954 millimeters.

1 liter = 35.197 fluid ounces.

1 ounce Avoirdupois = 28.34954 grams.

1 gram = 15.43235 grains.

LABORATORY EQUIPMENT.

Student's Individual Equipment with approximate cost.

171.	1 Beaker, No. 2 Griffin's, with lip, Bohemian glass	$ .10
	1 Long right angle bend, 8 cm. × 45 cm.	.02
	3 Medium right angle bends, 8 cm. × 15 cm.	.06
	3 Short right angle bends, 8 cm. × 8 cm.	.04
	2 500 c.c. bottles, 3.5 cm. across the mouth	.10
	6 5 oz. flint glass quinine bottles, 3.5 cm. across the mouth. Whitall Tatum Co.	.18
	1 Combustion spoon, ½ inch	.15
	1 Royal Berlin porcelain evaporating dish, 3 inch.	.25
	1 File, three cornered, 4 inch	.10
	1 Carbon dioxide flask, holding 285 c.c. when filled to the brim, and measuring 3.5 cm. across the mouth	.15
	1 Pair of iron forceps, 4.5 inch. Eimer & Amend new catalogue, No. 3335	.09
	1 Funnel of glass, 2½ inch	.05
	1 Funnel tube, 18 inch	.10
	8 Squares of glass, two inches square	.08
	1 Iron wire loop. No. 18 standard wire gauge	.01
	1 Mortar and pestle, E. & A., No. 4093, No. 7 2½ inch	.25
	1 Glass stirring rod	.01
	¼ lb. soft German glass tubing, inside diam., 5½ mm., outside diam., 7½ mm.	.10
	3 2½ inch rubber connectors, made from ¼ inch rubber tubing, double thick, white stock	.06
	1 8 inch piece of $^3/_{16}$ inch black rubber tubing of pure gum, double thick	.08
	3 No. 8 two-hole rubber stoppers, small diam., 33 mm.	.45
	1 No. 3 one-hole rubber stopper	.05
	12 Test tubes, 6 inch × ¾ inch	.24
	1 Wire test tube holder	.10
	1 Wooden test tube rack	.25
	3 Pieces of copper wire	.01

Equipment for general use.

One set of the following material should be constantly available at every laboratory position:

172.		1000 c.c. wide mouth bottle	$ .10
	1	bulb test tube. Eimer & Amend, No. 4877. These tubes are a little too thick on the bottom and are much improved by having a second small bulb blown on the end	.10
	2	Bunsen burners with tubing	.70
	1	Foot rule with centimeter scale	.10
	1	Square of cobalt glass, 4″ × 4″	.10
	1	500 c.c. filter flask with tubing	.20
	1	Package of cut filter paper, 10 cm. diam.	.10
	1	Brass filter pump, E. & A., 3250	1.25
		Filter paper discs. These may best be cut from ordinary sheet filter paper with a steel punch.	
	1	50 c.c. graduated cylinder, double graduation	.50
	1	2 inch iron crucible, E. & A., 2879	.30
	1	2 inch square iron plate with depression in center	.05
	1	Lamp stand of iron. Rod 18 inches high and 3/8 inch diam.	.35
	1	3 inch ring for stand	.05
	1	2 inch ring for stand	.05
	1	Clamp for stand, E. & A., 2750	.40
	1	Package of wrapping paper	.05
	1	2 inch porcelain sieve funnel, E. & A., No. 3357	.50
	1	Rubber stopper for funnel	.10
	1	Litmus paper bottle	.15
	1	8 oz. Conc. sulphuric acid bottle	.23
	1	" Dil. " " "	.23
	1	" Conc. nitric acid bottle	.23
	1	" Dil. " " "	.23
	1	" Conc. hydrochloric acid bottle	.23
	1	" Dil. " " "	.23
	1	" Ammonium hydroxide bottle	.23
	1	" Sodium hydroxide bottle	.23

These bottles have glazed white labels with transparent letters and formulas, Whitall Tatum Company.

1 Tin cup for shelf for water pan, with hole in bottom and side.......................... $ .05
1 Common centigrade thermometer, E. & A., 4894 graduated up to 130° C. .... .50
1 Bundle of tooth picks for splints..............
1 Water pan, any kind of a rectangular pan over 4 inches deep............................. .15
1 Iron wire gauze with asbestos center, E. & A., 7262 .... .12
1 Voltameter, see special fig. 4, made to order by Bausch & Lomb, duty free, glass part only... 2.00
1 Harvard trip scale, E. & A., 2142.............. 7.00
1 Set of weights, E. & A., 2204, 1 g. to 500 g. ... 2.25
1 Test tube cleaner with brush on end.......... .10

Additional Apparatus of Use in Making up Solutions and in Making Preparations.

Hot plate $5.00
2 8 liter flasks............................ 1.50
2 2½ liter flasks.......................... .80
1 12 inch Royal Berlin porcelain evaporating dish 2.00
1 5 inch glass funnel.......................... .30
1 Royal Berlin porcelain casserole, No. 6......... 1.50
1 6 inch porcelain sieve funnel, E. & A., 3356.... 2.00
6 No. 6 Griffin's beakers with lip................ 1.20
Folded filter paper, 38 cm. ....
1 Sodium press 12.00

Reagents sufficient to enable a class of twenty students to perform the experiments in this book, with approximate cost.

173. Alcohol, 1 pint.............................. $ .35
Potassium alum, ½ lb. .... .10
Aluminum sulphate, 1 lb., c.p. .... .25
Aluminum wire, 1 oz. .... .10
Aluminum foil, 1 oz. .... .10
Ammonium chloride, ½ lb. .... .12
Ammonium nitrate, ½ lb. .... .25
Ammonium thiocyanate, 1 oz. .... .05
Ammonium sulphate, com., 2 lb. .... .20

Ammonium sulphide, made by passing hydrogen sulphide into ammonium hydroxide until saturated and then adding an equal volume of ammonium hydroxide.

Arsenic, metal, ½ lb.	$ .30
Arsenic trioxide, ¼ lb.	.10
Borax, 1 lb.	.10

Bromine water, made by shaking liquid bromine with water and keeping an excess of bromine in bottle.

Bromine, 2 oz.	.25
Barium chloride, c.p., 2 oz.	.05
Barium nitrate, com., ¼ lb.	.10
Cadmium sulphate, 1 oz.	.20
Calcium chloride, com., ½ lb.	.05
Calcium nitrate, com., 1 oz.	.15
Calcium oxide, lime, 2 lb.	.05
Carbon disulphide, 2 lb.	.44
Charcoal, animal, 1 lb.	.02
Charcoal, wood, 1 lb.	.01
Chloride of lime, 1 lb.	.05
Chlorine water	
Chrome alum, ¼ lb.	.05
Coal, bituminous, 3 oz.	
Copper, sheet, 1 lb.	.40
Copper wire, No. 18 standard wire gauge, ¼ lb.	10
Copper gauze, 80 mesh, 10 grams	20
Copper scale, ¼ lb.	.10
Copper sulphate, com., 1 lb.	.10
Cream of tartar, 1 lb.	.35
Ferrous sulphate, com., 2 lb.	.06
Ferrous sulphide, com., lumps, 5 lb.	.50
Grape sugar, 1 oz.	.02
Hydrofluoric acid, 4 oz.	.50
Iron filings, ½ lb.	.05
Iron wire, fine, 2 lb.	.10
Iron wire, 2 lb. No. 16 standard wire gauge	.20
Iodine, 1 oz.	.25
Lead wire, 1 oz.	.05
Lead nitrate, com., 1 oz.	.02

Lead oxide, litharge, ¼ lb.	$ .03
Lead oxide, red lead, 1 oz.	.03
Magnesium wire, 1 oz.	.55
Magnesium oxide, ½ lb.	.25
Manganese dioxide, 1 lb., powder	.08
Marble, 1 lb.	.10
Mercury, 1 oz.	.10
Mercuric chloride, 2 oz.	.20
Mercuric oxide, red, 2 oz.	.20
Oxygen mixture, made by mixing two parts of pulverized potassium chlorate with one part of powdered manganese dioxide, 1 lb.	.30
Phosphorus, red, 1 oz.	.10
Phosphorus, yellow, 1 oz.	.10
This phosphorus should be granulated as described under Laboratory Suggestions. **176.**	
Potassium bromide, c.p., 2 oz.	.20
chloride, c.p., ½ lb.	.15
chlorate, c.p., ½ lb.	.10
chromate, 1 oz.	.05
dichromate, ½ lb.	.07
ferrocyanide, c.p., 1 oz.	.02
ferricyanide, c.p., 1 oz.	.05
hydroxide, pure by alcohol in sticks, 1 oz.	.10
iodide, c.p., 1 oz.	.30
nitrate, com., 1 lb.	.05
permanganate, 1 oz.	.05
sulphate, 1 oz.	.05
tartrate, normal salt, ½ lb.	.40
Quinine sulphate, ¼ oz.	.10
Sand, 1 oz.	
Silver nitrate, 1 oz.	.50
Silver wire or foil, 6 g.	.20
Sodium bicarbonate, c.p., 1 lb.	.20
carbonate, soda ash, 2 lbs.	.10
chloride, table salt, 2 lbs.	.05
nitrate, com., 1 lb.	.10
nitrite, com., 1 lb.	.10

Sodium phosphate, common cryst., 1 lb.	$ .10
sulphate, com. Glauber salt, 1 lb.	.03
thiosulphate, 1 lb.	.05
Sodium, metal, 1 oz.	.25
Sodium wire,—this is made when needed by means of a sodium press.	
Starch, corn starch, 1 lb.	.10
Stannous chloride, made when needed by dissolving tin in hydrochloric acid.	
Strontium nitrate, com., ½ lb.	.07
Sulphur, brimstone, 3 lbs.	.15
Tartaric acid, ½ lb.	.15
Tin, granulated, ¼ lb.	.15
Water glass, ½ lb.	.05
Zinc, granulated, 2 lb.	.40
oxide, ¼ lb.	.03
sulphate, ¼ lb.	.02

Stock Acids and Solutions for a class of twenty, for which it is desirable to have special reagent bottles.

174.	Sulphuric acid, conc., c.p., specific gravity 1.84, cont. 98% acid, 10 lbs.	$ .80
	Sulphuric acid, dil., made by mixing one part of the conc. acid with three parts of distilled water. s.g. 1.28, cont. 38.13% acid.	
	Nitric acid, conc., c.p., s.g. 1.42, 8 lbs.	.80
	Nitric acid, dil., made by mixing the conc. acid with an equal volume of water, s.g. 1.28, cont. 45.55% acid.	
	Hydrochloric acid, conc., c.p., s.g. 1.2, cont. 40% acid, 8 lbs.	.80
	Hydrochloric acid, dil., made by mixing equal volumes of the conc. acid and distilled water, s.g. 1.1, cont. 19.8% acid.	
	Ammonium hydroxide, conc., c.p., s.g. .90, 2 lbs.	.20
	Ammonium hydroxide, dil., made by mixing one volume of conc. ammonia with three vols. of water, s.g. .97 and cont. 15% ammonium hydroxide.	

Sodium hydroxide solution, conc., made by dissolving 1000 g. commercial granulated sodium hydroxide in 1000 c.c. of water in an iron dish, allowing the solution to settle for a week and siphoning off the clear solution; s.g. 1.51 and cont. 51% sodium hydroxide. 2 lbs. crude NaOH .... $ .20

Sodium hydroxide solution, dil., made by diluting one volume of the conc. solution with three vols. of water; s.g. 1.18 and cont. 16% NaOH.

Alcohol, ethyl, 95%.

Ammonium sulphide.

Barium chloride, 10% solution.

Bromine water, saturated with bromine.

Calcium hydroxide, saturated solution made when wanted.

Carbon disulphide.

Chlorine water. Made when wanted.

Copper sulphate, 10% solution.

Lead nitrate, 10% solution.

Litmus solution, made by dissolving about 1 g. of azolitmin in two liters of distilled water and filtering.

Mercuric chloride, 5% solution.

Potassium chromate, 10% solution.

Potassium dichromate, 10% solution.

Potassium ferrocyanide, 10% solution.

Potassium iodide, 5% solution.

Potassium permanganate, 10% solution.

Silver nitrate, 5% solution.

Sodium phosphate, 10% solution.

Stannous chloride, 10 g. of tin to 100 c.c. of acid.

Suggestions for the Laboratory.

175. **Hoods.** There should be one hood for every four students in a laboratory class.

Lockers. Each student should have an individual locker with a combination lock.

Distribution of Gases. If possible there should be an independent pipe system with an outlet at each position so arranged that any gas such as oxygen, hydrogen, carbon dioxide or air could be delivered at a moment's notice.

Ordinary ¼-inch gas pipe is large enough for the purpose.

Oxygen. This should be made as wanted and kept in a large gas holder.

Hydrogen. Should be furnished by a large Parsons' generator. E. & A. 3611. This generator has been used in this laboratory for six years, and has given perfect satisfaction. It has run for one year without any attention whatever, the gas being used intermittently.

Carbon Dioxide. This gas might well be made in a Parsons' generator, but it is much cheaper to buy it in liquid form in large cylinders and distribute it through a reducing valve.

Hydrogen Sulphide. A Parsons' generator should be set up in the basement or attic and connected with each hood by means of lead or iron pipe. It requires little or no attention during the year. At the most a mere refilling of the reservoir.

Sulphur Dioxide. This should be bought in liquid form in tin cylinders holding about five pounds. But to avoid accidents due to defective cylinders the whole should be inclosed in a piece of iron pipe with a screw cap on each end, through one of which the brass cock of the cylinder projects.

Chemicals. These should be pure as far as possible. It does not pay to use impure acids. Buy all chemical supplies direct from the manufacturer. Baker & Adamson, of Easton, Pa., furnish very satisfactory chemicals at very reasonable prices.

176. **Granulated Yellow Phosphorus.** This cannot be bought, but may be readily made in the laboratory as follows:

Stand a nine liter flask having a cylindrical neck at least two inches in diameter in the sink, and fill it with cold water up to the beginning of the neck.

Contract the end of a one-inch carbon filter tube by heating, to a one-eighth inch hole. Push this end barely through the center of a disc of sheet cork just a little larger than the diameter of the neck of the flask. Now push the cork in a horizontal position down into the neck of the flask until the filter tube is entirely within the neck, the top of the tube being about an inch below the top of the flask. Add cold water through the tube until the surface reaches the cork, and fill the neck with boiling water. Cut up all the stick phosphorus into pieces not more than one inch long, and quickly drop them one at a time into the filter tube. They will melt almost at once and run down into the flask in the form of shot. The first to go through will probably go into the undercooled condition and operations must be suspended until they crystallize. Add boiling water to the neck of the flask from time to time. A pound of phosphorus can be made into thousands of pellets in this way in less than half an hour. Reasonable caution must be used in handling so much phosphorus.

Bunsen Burner Igniter. A single wire circuit should be run to a pointed brass contact piece on the shelf in front of every Bunsen burner in the laboratory. A ten-cell dry battery and spark coil are put in the circuit, and the other end connected to the gas pipe system. A loose bare copper wire is run through every rubber tube and projects into the gas nipple and burner nipple.

It is only necessary to turn on the gas and touch the top of the burner to the contact piece to get a light.

This scheme has been in use in this laboratory for about four years and is perfectly satisfactory.

It would be a still better plan to run a two-wire system and use a double contact piece, the circuit being closed by the tip of the burner. If this were done, the 110 volt current could be turned on through two 32 candle power lamps abreast and a much hotter spark would be obtained.

Suction Pump. Every position should be provided with a suction pump and a filter flask with a small porcelain sieve funnel.

Master Key for Lockers. Have each combination lock engage with a common lock set in the jamb. All the common

locks being alike, one key in the hands of the instructor will enable him to open any locker in the laboratory instantly.

Glass Tubing. This should be of the soft German variety, and the standard size for all glass tubes used in the experiments should be 7½ mm. outside diam. and 5 mm. inside diam.

Dry Reagents. All dry reagents that are regularly issued to the students and are not altered by exposure to the air should be kept in 5-oz. wide-mouth bottles, as many of each kind as there are positions. And when issued there should be a small horn spoon in each bottle.

Horn Spoons. These are double horn spoons 5 inches long, E. & A. No. 4706. They cost about six cents apiece, duty free.

Flasks and Bottles. All flasks and bottles should be selected with the same sized openings in order that only one size of rubber stopper need be used.

Care of Wooden Tables.

Solution A. Shake up Bismark brown with wood alcohol in a gallon bottle. 100 grams of the color to a quart of alcohol.

Solution B. A solution of orange shellac, best obtained at the painter's.

Solution C. Made by mixing equal volumes of A and B.

Solution D. Dissolve one or two ounces of beeswax in one quart of boiled linseed oil with the aid of heat in an iron dish or pot. Make this up to a gallon with turpentine.

Paint the tables all over with solution C and from time to time in spots where they need it. About once a month rub them all over with the oil wax polish, solution D.

The Bismark brown is quickly turned to a light color by alkalies but the color may be quickly restored by rubbing the spot with dilute nitric acid, half water and half strong acid. It is a good plan to rub them all over once in a while with the dilute nitric acid. This changes the oil to a very hard hornlike substance and makes them look as though they had just been revarnished.

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